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PRINCIPLES OF CHEMISTRY APPLIED TO THE HOUSEHOLD

Principles of Chemistry
Applied to
The Household

An Elementary Text-Book

BY
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BOSTON
THE BOSTON COOKING SCHOOL MAGAZINE CO.
1919

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The Boston Cooking School Magazine Co.

The University Press, Cambridge, U. S. A.

PREFACE

THE writing of this text was undertaken because the authors felt the need of such a book in some of their own classes. In addition to college entrance work, short courses in general chemistry were given — three or four forty-minute periods a week throughout the school year — in which there was no need to meet requirements of college entrance examinations, and in which it was felt desirable to emphasize the applications of chemistry to practical things of interest to the students. Several excellent text-books were found which contained much general information in such matters, and some laboratory guides which suggested interesting experiments on foods and cleaning; but there was no text-book that contained just the desired relation between these features and the simple, fundamental work which it was felt should precede such applications. This book was written to fill such a need. It has been tested in class-room and laboratory and found to be practical. Its preparation has helped the authors very much in stimulating and unifying their class-room work. If its use by others be of any like help, their decision to publish it will have been justified.

The advantage of the first half of the book is felt to be its unity of treatment, centering around the idea of chemical change, a unity not possible where the attempt is made to write a descriptive book as well. Part I is brief and fundamental, and the ideas presented there are applied throughout Part II to the problems of food and cleaning. It is this sequence of thought that is felt to be the most valuable feature of the presentation. The chapters on theory have been made very simple, and have been directly applied, wherever possible, to the work that precedes and that follows them. A supplement is added containing general information on household substances, and additional laboratory experiments, which may be introduced wherever needed to adapt the work to the particular circumstances of the course offered. Standard texts and experiments have been freely used, and the many references have been carefully chosen to broaden the student's horizon still further.

Professor Dewey has said: "It seems to me that our present methods too largely put the cart before the horse; and that when we become aware of this mistake we are all too likely to cut the horse entirely loose from the cart, and let him browse around at random in the pastures without getting anywhere. What we need is to hitch the horse of concrete experience with daily occupation and surroundings to a cart loaded with specialized scientific knowledge. It is not the business of high-school science to pack the cart full — that will come later. It is its business to make such a good job of the hitching that every pupil who comes under its influence will always find in himself a tendency to turn his crude experience over into a more scientific form, and to translate the bare science he reads and hears back into the terms of his daily life. When we do this, we shall find, I am confident, the crop of scientific specialists increased, not diminished, while we shall have a citizenship of men and women really intelligent in judging the affairs of life."[1]

It is in the hope of improving the hitching, so to speak, that the principles of chemistry have been here applied to foods and cleaning processes rather than to technical laboratory experiments in qualitative analysis or to descriptive work in commercial chemistry. The hope is that such a course will prove more useful and more genuinely educative.

---

[1] Extract from address before the Science Section of the N. E. A., July, 1916.

While it has always been kept in mind that the book is one for beginners in chemistry and so should be simple in treatment, yet attention has been paid to the most recent investigations, such as the relation of proteins to amino-acids, and the subject of vitamines. The text has also had the great advantage of valuable criticism. The authors owe their thanks to Prof. Elmer P. Kohler of Harvard University for his criticism of the chapters on theory, to Prof. A. G. Woodman of the Massachusetts Institute of Technology for his help in reviewing the work on food, and to Prof. Charlotte A. Bragg of Wellesley College for a consideration of the entire manuscript.

Our thanks are also due to Miss Helen Janet Wright and Miss Emily Tucker for their help in preparing the manuscript for publication.

H. T. R.

Boston, January, 1918.

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PRINCIPLES OF CHEMISTRY APPLIED TO THE HOUSEHOLD

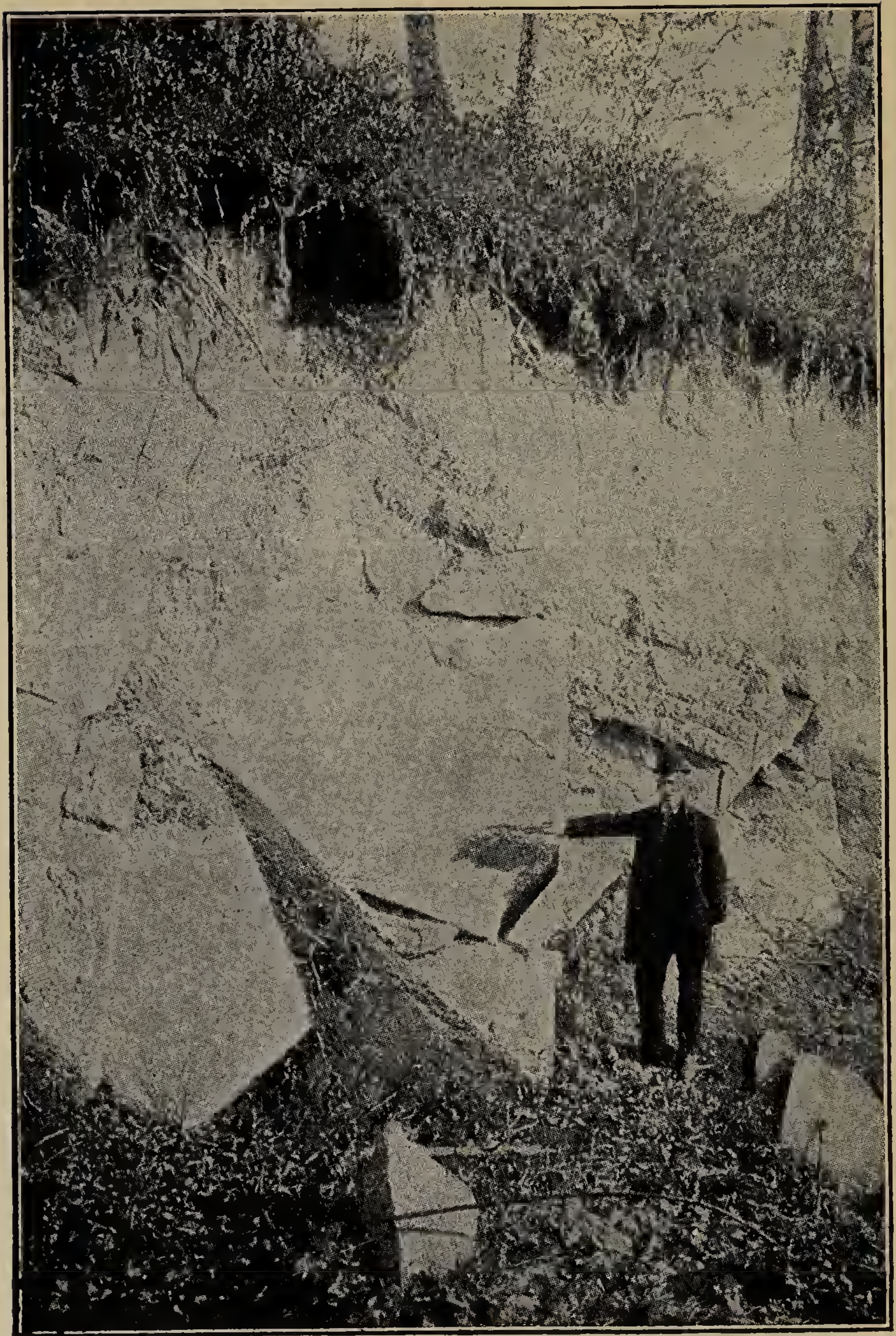
PART I

CHAPTER I

CHEMICAL CHANGE

WE cannot study successfully the applications of chemistry to the household without first learning the foundation principles of chemistry itself. The science is both old and new, — old in the sense that men studied the changes undergone by matter in the days of Greece and Rome, and even earlier, and new in the sense that only within the last hundred and fifty years have they discovered the truth about certain very important changes, as, for instance, what takes place when a substance burns. An understanding of such happenings, or phenomena, put chemists on the right road, and has led to a marvelous development of the science in recent times; while within the last twenty or thirty years new discoveries have led to new theories and new methods, so that we speak of modern chemistry as quite a different thing from the chemistry of the alchemist of the Middle Ages, or of the men who believed in "phlogiston," "light-matter" and "heat-matter" of the eighteenth century.

You will understand from what has been said that chemistry deals with the changes that take place in matter. Such changes as the weathering or crumbling of rocks, the coloring of the leaves in the autumn, the proc-



From "Rocks, Rock-Weathering and Soils"
By Geo. P. Merrill, published by The Macmillan Co.

Weathered Granite, District of Columbia

The exposed face of the rock shows the gradual change from hard, fresh granite at the base, through a middle section of looser friable material, to sand and gravel, and finally, soil at the top. Such changes are largely physical, but new substances are also formed by the action on the rock of the oxygen, carbon dioxide and water of the air.

esses of food digestion in both plant and animal bodies, the souring of milk, and so forth, are as much the subject matter of chemistry as are those changes that we can cause to happen by various means in our laboratories. In fact, modern chemistry has been able to imitate the processes of nature to some extent, and so to produce in the laboratory such substances as certain kinds of sugar, and vanillin, the flavoring found in the vanilla bean, which it was formerly thought could be made only by living plants or animals.

Moreover, by causing matter to undergo changes in the laboratory, and observing very carefully what happened, chemists have been able to discover much about the composition of matter, and the laws which it obeys, and to make theories which cannot exactly be proved but which help us to go on experimenting and discovering new truths, and getting a clearer understanding of our material world.

In this course we shall be, as far as possible, experimenters, and we shall begin by studying carefully some common changes that matter undergoes.

EXPERIMENT 1

Purpose.

To study changes in matter.

Directions.

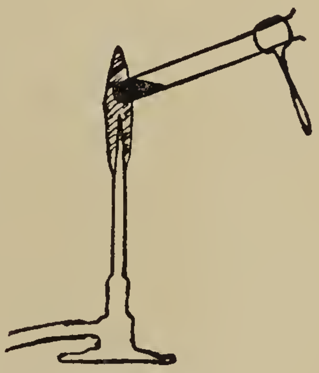


FIG. 1

a. Place a small quantity of sugar — about 3 c. c. — in a dry test-tube, and heat (Fig. 1). Notice how the sugar changes in appearance, and what products form. Are they solid, liquid or gaseous? (Fumes are solid matter in a very finely divided condition, carried upward by rising gases.) From time to time try to set fire to the fumes by putting the mouth of the test-tube at the flame. Describe what happens. Continue the heating until no further reaction occurs. Allow the tube to cool, and then wrap an old cloth around it and break it open. Describe the contents.

> **Questions.** Have new substances been formed in this experiment? Do you know of any way in which the sugar may be recovered? Would you call this change that the sugar has undergone a permanent or a temporary one?

NOTE. The black substance left in the test-tube is carbon, a substance which occurs in an impure state as coal and charcoal.

b. Dissolve some salt in water in a small beaker. Tell in what very evident way the salt has changed. Pour the solution into an evaporating dish and boil off the water over an asbestos wire gauze (Fig. **2**). Be careful not to crack the dish. Describe what remains in the dish. Taste it. What is it?

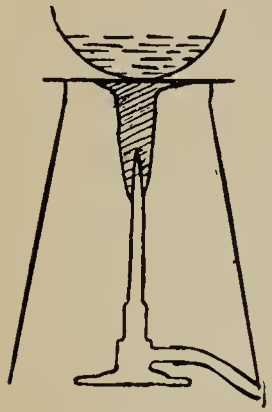


FIG. 2

> **Questions.** Was a new substance formed in this change? Was the change undergone by the salt when dissolved a permanent or a temporary one?

c. Place a few crystals of iodine in a test-tube, and heat (Fig. 1). Describe what happens. Try to pour out the purple vapor. Is it heavier or lighter than air? Heat only the end of the test-tube and continue until the end is empty. Describe what you see on the sides of the test-tube. What is this substance?

> **Questions.** Was a new substance formed? Did the iodine, in changing to a vapor, undergo a permanent or a temporary change?

NOTE. Such a change from solid to vapor is called subliming or sublimation. We say the iodine was sublimed.

d. Take a piece of magnesium ribbon about one and a half inches long and, holding it with the forceps, put it in the flame (Fig. 3). What happens? What remains? Feel the substance. Has this substance the properties of magnesium, that is, is it like magnesium in color, hardness, luster, etc.?

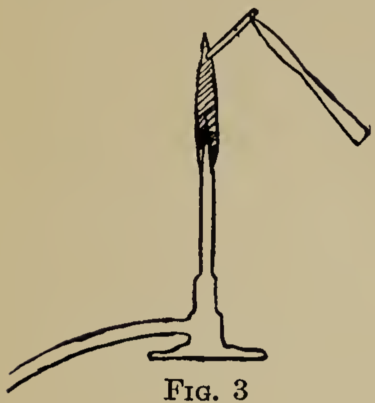


FIG. 3

Questions. Has a new substance been formed? Could you easily recover the magnesium? Has the magnesium undergone a permanent or a temporary change?

e. Place a rubber rod near some light object like a pith ball or a few scraps of paper (Fig. 4). Does it affect the object? Now rub the rod hard and rapidly with fur or woolen cloth for a few moments, and again bring it near the pith ball. Result? Try the effect on the pith ball ten minutes after rubbing? Result?

Questions. Was the rubber rod changed, if so, in what way was the change shown? Has a new substance been formed or is it still rubber? Was the change temporary or permanent?

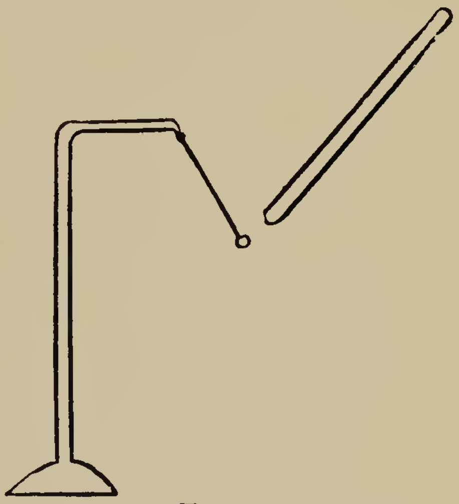


FIG. 4

NOTE. We say the rubber rod was electrified or was given a charge of electricity by friction or rubbing, but we do not know just what this change is.

f. Place a piece of ice in an evaporating dish and heat on the asbestos wire gauze (Fig. 2). Result? Continue heating until the dish is empty. Think hard about this.

Questions. What two changes occurred? Are ice, water and steam different substances, or may one be changed directly into another? Can these changes be reversed easily, that is, can steam be changed to water and water to ice? Are such changes temporary or permanent?

If we sum up what we have learned by these experiments, we shall come to the conclusion that the charring or burning of the sugar and the burning of magnesium were alike in that when they were ended we could not, by any simple treatment of the new substance or substances left, reproduce the thing with which we started. These changes are permanent and result in the formation of new substances which do not return to sugar and magnesium no matter how long we leave them or how much we alter the conditions affecting them, such as temperature, for instance. If we consider the other changes, the dissolving of salt in water, the subliming of iodine, the electrifying of the rubber rod, the melting of ice and the evaporation of water, we shall see that these are different from the other two. The substances may have been given new properties; for example, the rubber rod acts differently after having been rubbed. In other cases they may have changed so as to seem like new substances: ice, water and steam are very different from each other, and iodine vapor is different from solid iodine. They may even have disappeared from sight entirely, as the salt did on dissolving. But in all these instances the original substance, in its original state and with its original properties, can be obtained again, either by waiting long enough or by changing the temperature of the new form. Steam if cooled will change to water and water to ice; iodine vapor, when it meets the cool sides of the test-tube, changes back into solid iodine; the rubber rod ceases to attract the pith ball after a time; and even the salt solution, if left long enough, loses its water by evaporation and we have again the original salt.

It is such changes as those undergone by the sugar and the magnesium that we study in chemistry; changes that result in new substances in place of those with which we start. Such changes usually take place rapidly only at higher temperatures than that of the air, and so we often use the heat of the Bunsen flame to raise the substances to the temperature at which they will change. There are other ways of bringing about such changes, which we shall use later.

DEFINITIONS

Substances are different kinds of matter, such as wood, brass, water.

Chemistry is the study of the substances that make up matter, and of the changes that they undergo.

A **chemical change** is one that results permanently in the formation of a new substance or substances.

QUESTION

Which of these phenomena are chemical changes: turning on of the electric light; souring of milk; changing color of leaves in the autumn; weathering of rock; breaking of glass; vaporizing of iodine; decay of apples; rusting of iron; freezing of water; burning of paper; striking a match?

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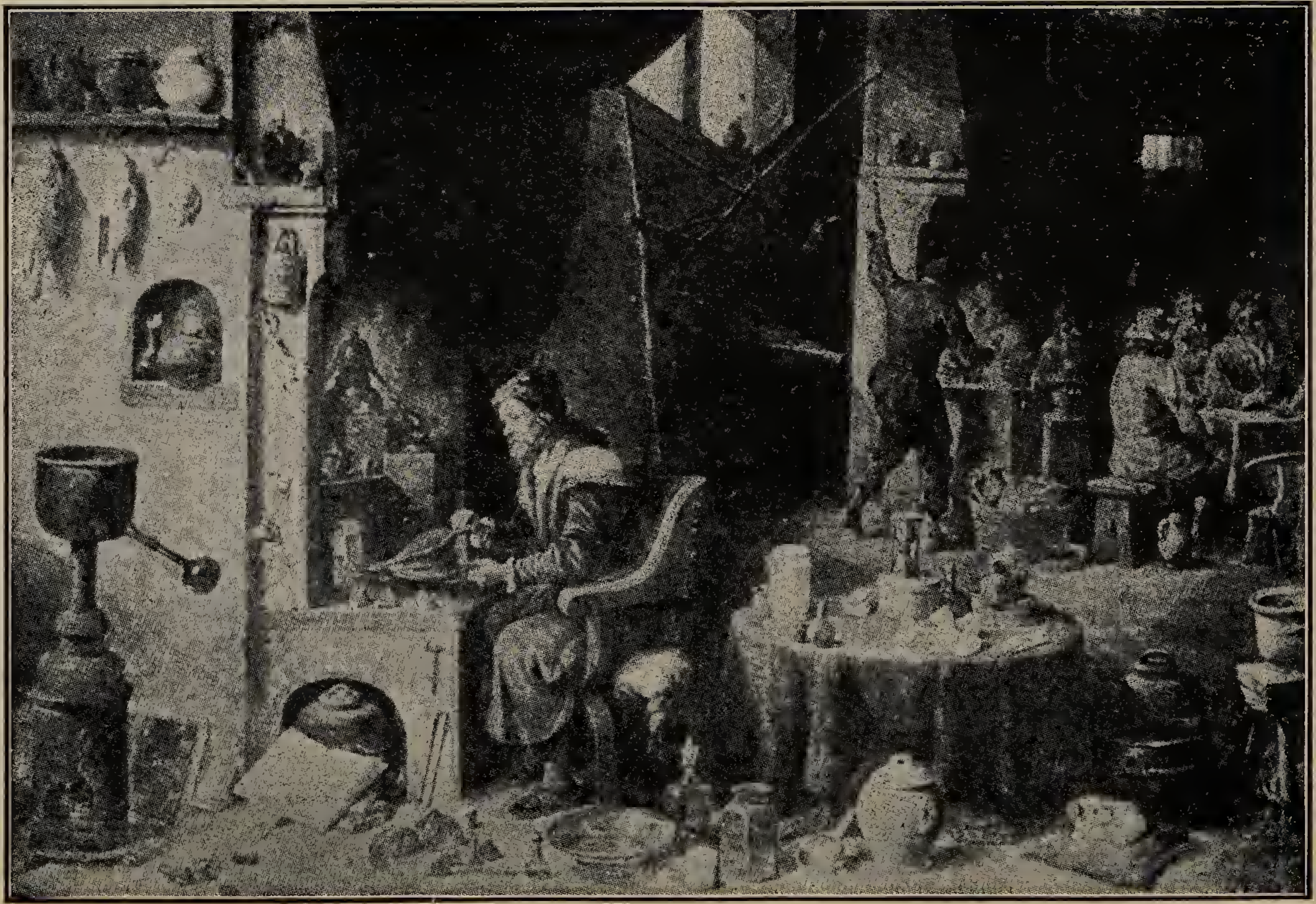


An Alchemist's Laboratory

Notice the rich hangings, the general air of disorder, the absence of balances. The apprentice in the background is operating the picturesque bellows which was used before such appliances as the Bunsen burner and the oxy-hydrogen blowpipe. The hour-glass marks the time.

CHAPTER II

SYNTHESIS

ONE entire class of chemical changes may be summed up under the head of burning. Many substances when heated go through this process. The experimenters of the eighteenth century, particularly Lavoisier in France and Priestley in England, were much interested in burning and performed many experiments to determine just what happened when a substance burned. The theory commonly believed at that time was called the phlogiston theory, and set forth that burning meant the driving out from a substance of something that they called phlogiston. As we know, many things, such as stones, will not burn even when heated very hot. These, it was thought, did not contain any phlogiston. We shall endeavor to study burning by heating in the air certain substances, mostly metals, and noticing carefully what happens to them. Before beginning, we may consider what the possible changes might be.

First, the substance burning might simply be changing over into another substance, just as the philosopher's stone of the alchemist's belief was supposed to change over baser metals like lead into precious metals like gold, or as water changes into steam. Second, the substance might lose something to the air, phlogiston or something else. Third, the substance might take in something from the air. It is plain to us that a very simple test will determine which of these three things is happening. In the first instance the substance will remain of the same weight, in the second instance it will lose weight, in the third instance it will gain weight.

This is evident to us, but the early chemists had not yet realized the necessity for using the balance. The importance of quantitative work was not understood until the fact now called the law of the conservation of matter had been established. This fact is that, regardless of the changes that we may cause it to undergo, we cannot create or destroy matter. This principle, now accepted as the foundation of all scientific work, gives value to such experiments as we are about to perform. Evidently by weighing our substance first, then heating it, and then reweighing, we may determine what has happened.

EXPERIMENT 2

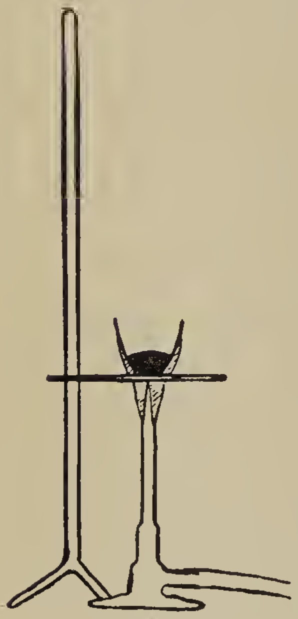


Fig. 5

Purpose.

To study the weight change on heating a metal.

Directions.

Weigh carefully a small porcelain crucible and its cover. Place in it a small quantity — about one fourth the capacity of the crucible — of the metal to be heated, weigh again, and place on a pipe-stem triangle (Fig. 5). Heat gradually, increasing the heat until the hottest flame is used. Describe any change which takes place in the appearance of the metal. Allow to cool with the cover on. Reweigh. Record results thus:

Data

Weight of crucible and cover . . . .
Weight of crucible, cover and metal
(give name) . . . .

Weight after heating — minutes . . . .
Change in weight (give gain or loss) . . ..

a. Copper filings. Free these first from grease by shaking vigorously in a small bottle with ether. Rinse thoroughly with fresh ether, and leave to dry between filter papers. Leave the cover off during heating. Use the hottest flame twenty minutes.

b. Iron filings. Hottest flame ten minutes, cover off.

c. Zinc dust. Hottest flame ten minutes, cover on. Lift the cover for a moment from time to time.

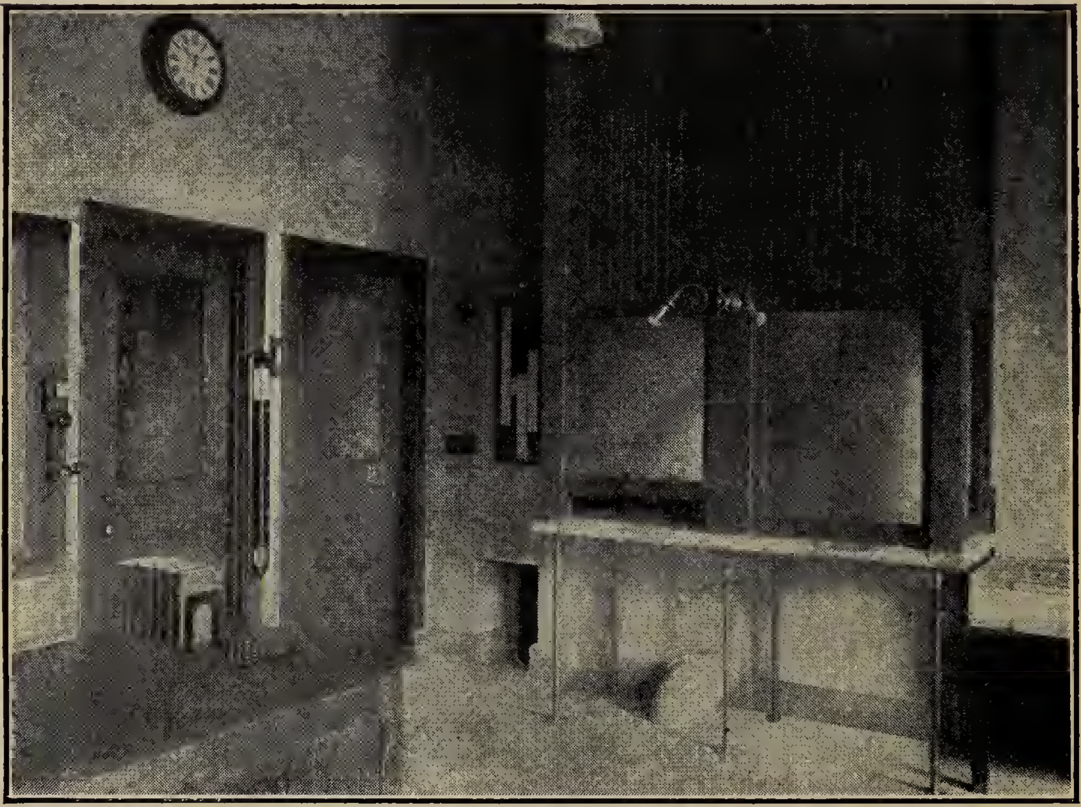
d. Granulated tin. Hottest flame ten minutes, cover off, stirring often with iron wire.

e. Powdered magnesium. Hottest flame ten minutes, cover on. Lift the cover for a moment from time to time.

Question. Assuming that crucible and cover remain the same in weight, how can you account for the change?

It is plain from all these instances that something has combined with the burning substance to increase its weight. That something must have come from the air. It is found that if the metal is heated in a closed vessel containing air, no increase of weight occurs. The new substance is formed, however, and when the vessel is opened air rushes in and the weight of the whole is increased. By these experiments in heating metals we have actually discovered anew for ourselves a truth that marked a turning point in the history of chemistry a hundred and fifty years ago. The men of those days formed a theory of burning, the phlogiston theory, and because their experiments did not seem to bear out that theory they considered, not that the theory was wrong but that the experiments were discredited. The experience of those early chemists should make us realize, as it finally did them, the necessity of

One end of a modern laboratory, showing the weighing room and balances.



Another corner of the same laboratory, showing the hood for vapors, and other necessary equipment of a modern laboratory.

proving our theories by experimenting, and of discarding them when our experiments prove them mistaken.

Burning, or the union of the substance in the air with the burning substance, is an illustration of a kind of chemical change called synthesis. This means a change producing a new substance by the chemical union of two or more substances. If we think about it we shall see how remarkable and wonderful a thing this is. We take a red metal like copper, and by causing it to unite with an invisible something in the air, we produce a black substance unlike both copper and air. We may take sulphur, a crumbly yellow solid, and by burning it cause it to disappear entirely. Where has it gone? Notice the odor that accompanies the burning, and you will have a clue to the answer. The product formed is an invisible gas with that characteristic, pungent odor. The sulphur has disappeared from sight, but still exists as part of the gas.

Such changes are extremely remarkable. What the cause is that makes the sulphur and oxygen combine we do not know. We say that they have chemical affinity for each other; but beyond very recent indications that an explanation of chemical affinity may be closely connected with electricity, we have no clue to the fundamental cause of chemical union. Chemistry has theories of the composition of matter and the ways in which matter acts which explain very clearly the circumstances of such phenomena, but we never really cease to wonder at them. These theories we shall study later. At present we shall go on from synthesis, or the building up of substances from simpler ones, to a consideration of other kinds of change.

DEFINITIONS

Synthesis is the building up of substances from simpler ones by means of chemical action.

A **law** is a generalized statement, founded on experimental work, to which there is no exception.

The **law of the conservation of matter** is such a statement, and sets forth that we cannot create or destroy matter.

QUESTIONS

1. Why was it so necessary to clean the copper filings from grease before heating them?

2. Why was the cover of the crucible left off during the heating of most of the metals? Why was it left on in the case of the zinc and magnesium powders?

3. Why was the tin stirred during heating?

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CHAPTER III

DECOMPOSITION

EXPERIMENT 3

Purpose.

To study decomposition.

Directions.

a. Weigh a hard glass test-tube, and place in it a quantity of the red powder given you, so that the weight is increased three or four grams. Weigh again accurately. Heat, at first gradually, then at a high temperature (Fig. 6). Describe the changes taking place. Introduce a spark on a joss stick into the tube. Result? When the powder has entirely disappeared, allow the tube to cool, and reweigh.

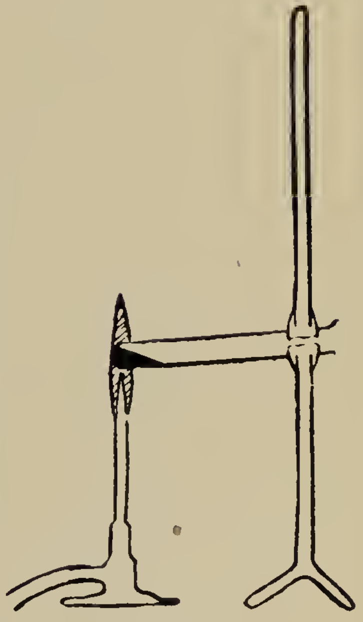


FIG. 6

Record thus:

Data

Weight of test-tube . . . .
Weight of test-tube and red powder . . . .
Weight after heating . . . .
Change in weight (give loss or gain) . . . .

Questions. What is the substance left in the tube? Was it formed by combining the red powder with something, or by driving something out of it? Give reason for your answer.

NOTE. This change is the reverse of those changes produced in Experiment 2.

b. Heat some potassium chlorate in a hard glass test-tube. Describe the changes undergone. Hold a spark on a joss stick in the tube. Result? Compare with the last experiment.

> **Question.** From the action of the spark do you think the potassium chlorate is combining with something or giving something off?

Both of the foregoing experiments are illustrations of change due to decomposition, or the breaking down of a substance into two or more simpler substances. The red powder evidently was made up of mercury and some invisible substance which went off into the air. This would suggest the idea that the powder might be made by burning mercury, and such is the case. The potassium chlorate when heated was decomposed into a different white solid and an invisible gas which caused the spark to burst into flame, or in other words made the joss stick burn more actively. We shall be justified in supposing that this was the same invisible gas as before.

Now we are going to decompose a very stable substance, water. Water may be heated very hot and so changed into steam, its gaseous form; and steam may be heated to very high temperatures without being decomposed. However, if a little sulphuric acid be put into the water, and then two electrodes introduced and a current of electricity passed through the water, the sulphuric acid helps the electric current in breaking up or decomposing the water and two gases result, one collecting at each electrode (Fig. 7). If a spark be introduced into the gas that has collected over the anode, or the electrode at which the current enters, the spark bursts into flame. Here is the same invisible gas again, this time found to be one of the two gases that, chemically combined, make up water. The gas over the cathode, which has collected twice as fast, is evidently a different gas, for when a spark is introduced the gas itself

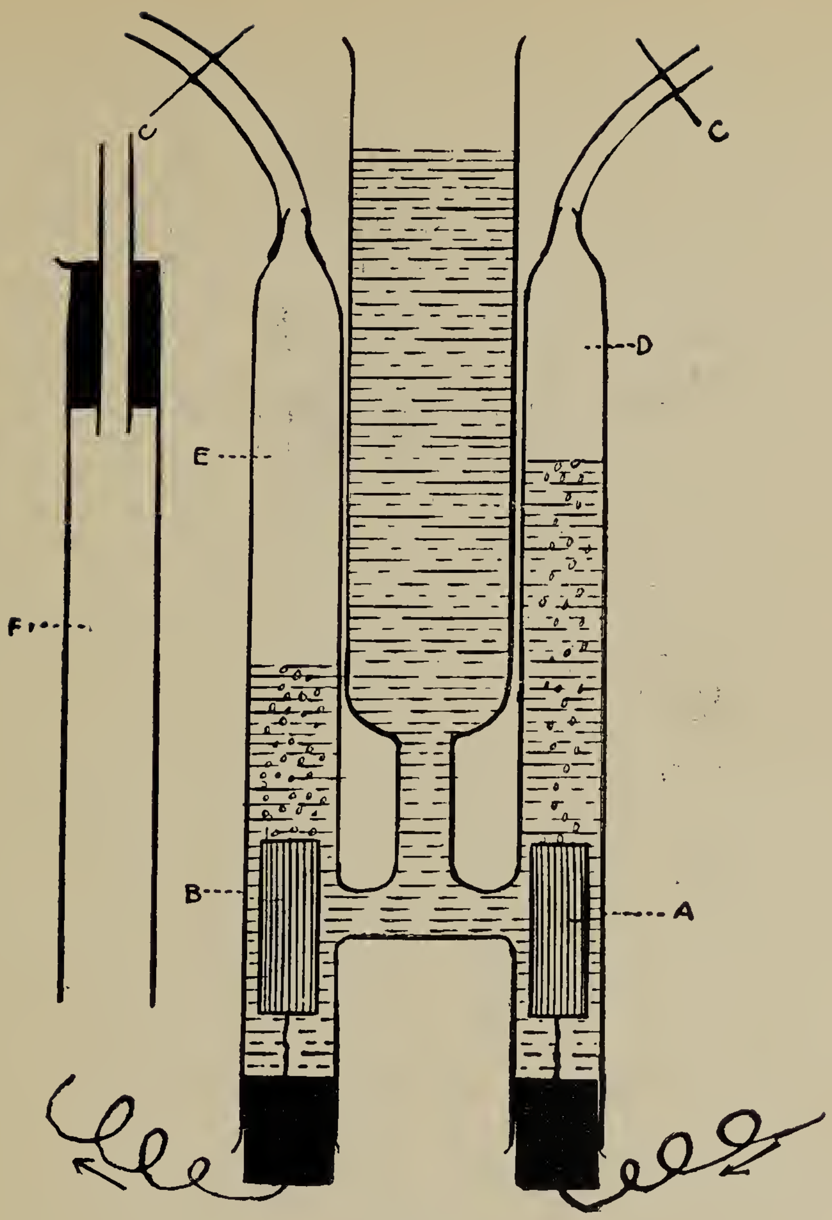


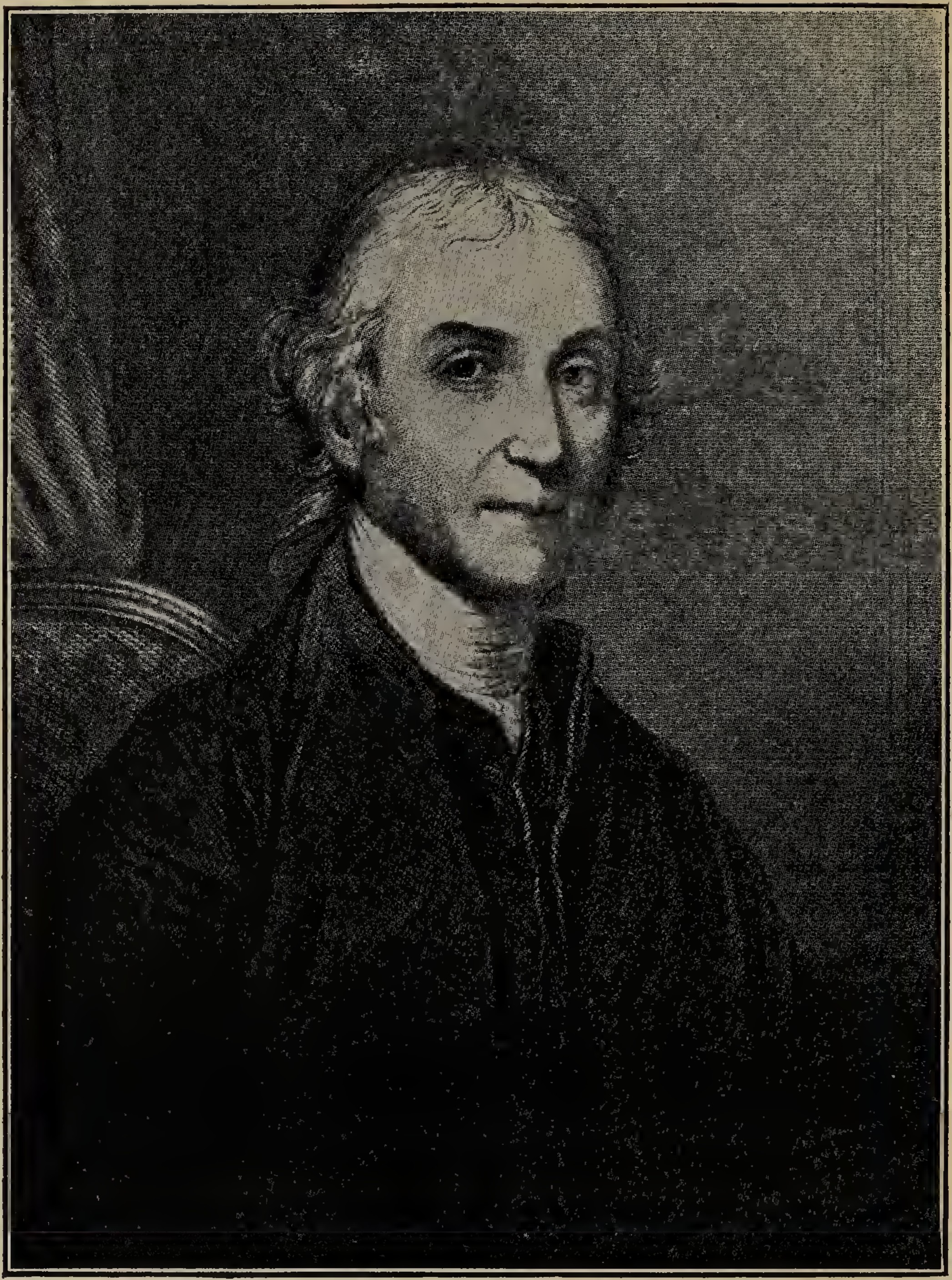
Fig. 7

The Electrolysis of Water

a. Anode, or electrode where current enters. *b.* Cathode, or electrode where current leaves. *c.* Pinchcocks on rubber tubing. *d.* Oxygen. *e.* Hydrogen. *f.* Collecting tube and attachment.

The gases may be collected easily after both tubes are full by using a piece of large glass tubing — three-quarters of an inch in diameter — and fitting into one end a one-holed stopper carrying a short piece of glass tubing which fits tightly into the rubber connection from the apparatus. Squeeze the rubber tubing to remove water, connect with the collecting tube, and place the thumb over the open end of the large glass tube. Then open the pinchcock, and at the same time loosen somewhat the pressure of the thumb. The gas will rush in, partly displacing the air, and the pinchcock is closed when the water reaches the rubber tubing. Remove the small tube from the rubber, and place the finger over it. Good tests for both hydrogen and oxygen may be obtained by this method.

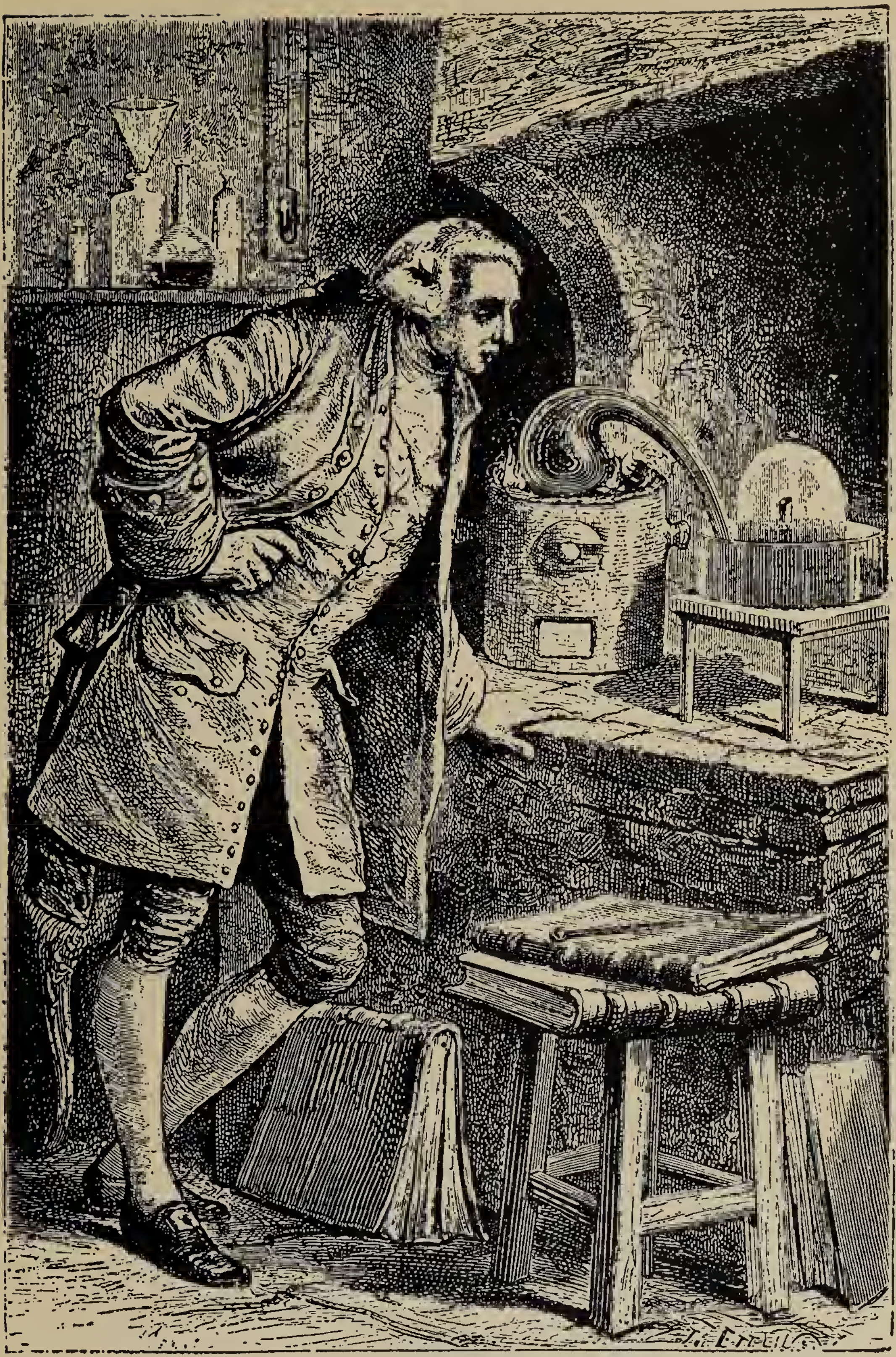
takes fire and burns with a pale blue flame. The latter gas is hydrogen, and the former is called oxygen.



The Perry Pictures

Joseph Priestley, 1733–1804

Lavoisier and Priestley, previously mentioned, are honored in this connection because they were the first to isolate this gas oxygen from the air, of which it is only a part, and to realize that it was the cause of burning. In the air, oxygen is found mixed with other gases, but not



Lavoisier Analyzing Air

The retort contained mercury, and opened into air confined over more mercury in the bell jar. When Lavoisier heated the retort, red powder formed on the mercury in it, and the air in the bell jar diminished about one-fifth in volume.

chemically combined with them because they are inert and do not unite readily with other substances. But oxygen is an active substance, and when many substances are heated to what is called their kindling temperature they enter into combination with the oxygen of the air. This is burning or oxidation and the products formed are called oxides. So water may be called hydrogen oxide, the red powder mercury oxide, and so on.

EXPERIMENT 4

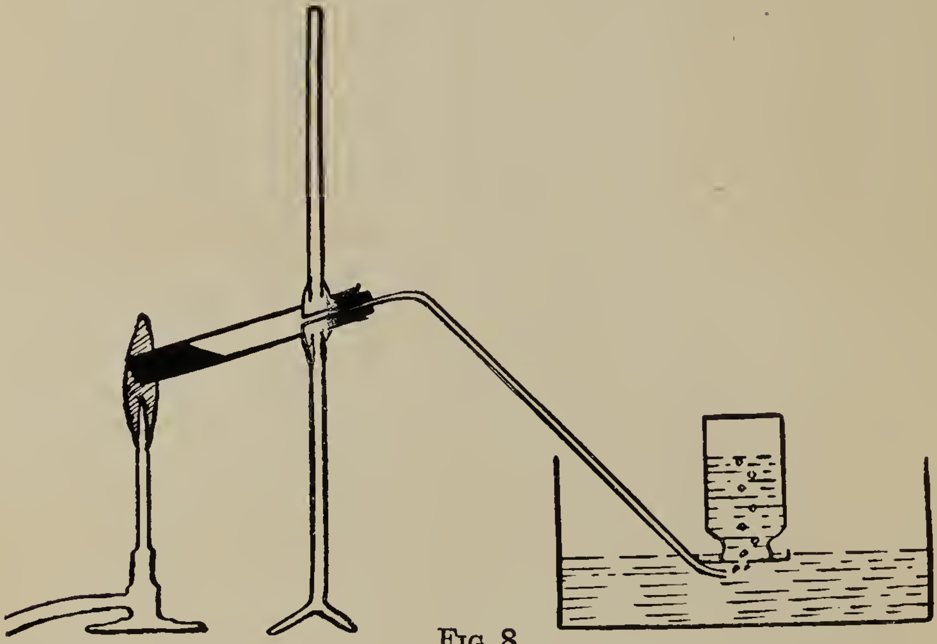


FIG. 8

Purpose.

To study the formation of oxides.

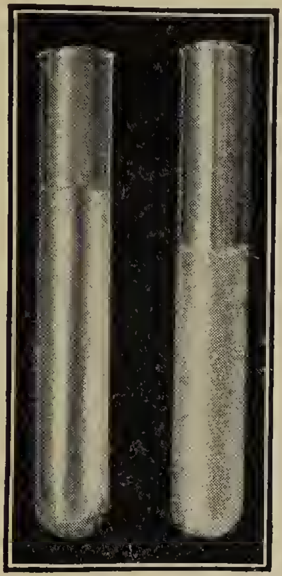
Directions.

Collect four or five bottles of oxygen by heating potassium chlorate (mixed with manganese dioxide to help the action) in a hard glass test-tube, and passing the gas through a delivery tube into bottles filled with water and inverted in a pan of water — a pneumatic trough (Fig. 8). Be careful not to remove the heat without first removing the end of the tube from the water.[1]

---

[1] The oxygen may conveniently be collected at the lecture table by the class, and each substance burned by one pupil, the others observing and helping.

Burn each of the following substances in a bottle of oxygen, describing the substance, its action when heated in air and in oxygen, and the nature of the oxide formed.



Clear and milky limewater.

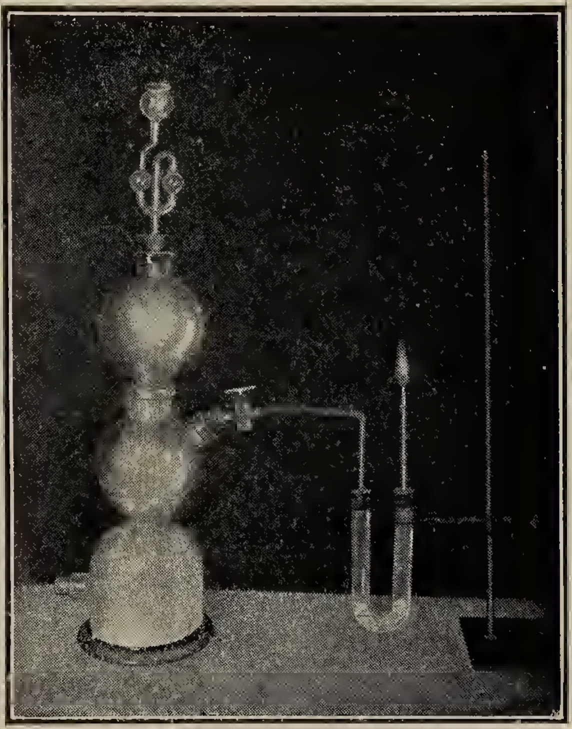
Iron burning in a bottle of oxygen.

a. Charcoal. Fasten a piece to the end of an iron wire, heat in the Bunsen flame until it glows, put quickly into the bottle of oxygen. After the action is over pour limewater into the bottle and shake. Result? This is a test for the gas carbon dioxide.

b. Red phosphorus. Heat in a deflagrating spoon until it is burning and then put into oxygen.

c. Sulphur. Follow directions for phosphorus.

d. Iron. Use picture wire. Dip the end first into melted sulphur until a sulphur tip is formed. Hold in the forceps, light



A jet of burning hydrogen, sometimes called the Philosopher's Candle.

the sulphur tip in the flame and then place in a bottle of oxygen.

> NOTE. The action of a substance, such as the manganese dioxide, which helps a reaction without taking part in it, is called catalysis.

DEFINITIONS

Decomposition is the breaking down of substances into simpler ones by means of chemical action.

Oxidation is the uniting of a substance with oxygen.

Burning is rapid oxidation usually accompanied by light and heat.

The **kindling temperature of a substance** is the temperature at which it will burn. This differs for different substances, and for the same substance under different conditions.

A **catalytic agent** is a substance which aids a reaction without directly taking part in it.

QUESTIONS

1. Judging by the electrolysis of water, what part by volume of water is oxygen?
2. Why do substances burn more vigorously in oxygen than in air?
3. Why was a sulphur tip made for the iron wire before burning it in oxygen? Why was picture wire used instead of a solid piece of iron?
4. Which has the higher kindling temperature, copper or magnesium? See Experiment 2.

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    Brownlee. Chemistry of Common Things, Chap X.
    Philip. Romance of Modern Chemistry, Chap. X.
    Philip. Achievements of Chemical Science, Chap. III.

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The Stick-Fire

CHAPTER IV

REVIEW — ELEMENTS AND COMPOUNDS

THE consideration of a common phenomenon, the burning of wood, will serve well to review what has already been learned about chemical change. Wood is made up of a number of very complex substances. When it burns it is first decomposed, by the heat applied to it, into simpler substances, as was the potassium chlorate. These may, in their turn, be further decomposed, and finally the simplest substances are oxidized by uniting with the oxygen of the air. Most of the substances formed are gases, like the products of burning carbon and sulphur, and these go off into the air. There is left a black mass which is charcoal, and recalls the substance left in the test-tube after the sugar had been heated. If this charcoal, or carbon, be heated strongly enough it will itself be oxidized, and only a few gray ashes left, which represent the part of the wood that cannot be burned. Burning of wood, therefore, illustrates both decomposition and synthesis. Most of the changes which take place in nature are complex, like this one, and involve several different types of change. For convenience they can be analyzed into simpler types.

Some natural changes, such as the decay of plants, the rotting of fruit, and the rusting of iron are essentially oxidations, although complicated by other factors. Here the processes take a long time to complete, but the products formed are oxides exactly similar to those resulting from rapid burning. In decay, certain bacteria present in the air play an important part. In the rusting of iron, the carbon dioxide and the moisture present in the air are necessary; but the product formed is, at least in part, exactly the same substance as that formed when iron was burned in oxygen. Spontaneous combustion, which starts sometimes where oily rags or waste are piled up in close places, is due in the beginning to a slow process of oxidation of the oil in the rags. If there is little ventilation the heat formed by this slow oxidation accumulates until the rags are hot enough to burn rapidly, or in other words have reached their kindling point, and so a fire is started.

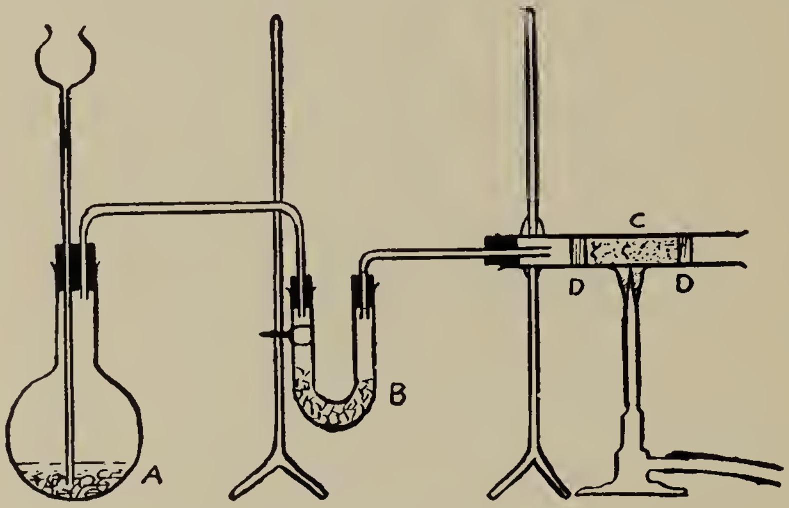

Fig. 9

Reduction of Copper Oxide by Hydrogen

a. Zinc and hydrochloric acid. *b.* Tube of calcium chloride which dries the hydrogen generated in *a.* *c.* Copper oxide. *d.* Asbestos plugs.

Just as one class of experiments can be grouped together under the heading oxidation, so the reverse of these changes all fall into one general class called reduction. If hydrogen be passed over copper oxide which is strongly heated, there is formed in the tube metallic copper, and steam issues from it (Fig. 9). The hydrogen has reduced the copper oxide to copper by combining with the oxygen in it to form water, or hydrogen oxide. Here reduction and oxidation are seen to be opposites, and occur simultaneously, the copper oxide being reduced and the hydrogen oxidized.

We have studied some types of chemical change, and learned the nature of some classes of substances. We have defined synthesis as the building up of substances from simpler ones, and decomposition as the breaking down of such substances into their simpler components. We have defined oxidation as one kind of synthesis and reduction as one kind of decomposition. We are now ready to make a study of the kinds of substances, and we shall group them into two main classes. These are elements and compounds. Elements are substances that we cannot decompose into simpler substances. Such are oxygen, carbon, sulphur, the metals, and many others. No way has yet been discovered by which they can be decomposed. Water was for a long time considered an element, because hydrogen and oxygen are so closely united to form it that it is extremely difficult to decompose. Since the common use of electricity it has been found possible to do this, and water is now known to be a compound, that is, a substance that can be decomposed into simpler substances. All the other oxides resulting from burning are of course compounds. There is another class of substances called mixtures, which are not elements or compounds, and which do not result from chemical union of elements. The elements or compounds making up mixtures have not passed through that mysterious change called chemical union, but are simply mixed together much as sugar and sand might be mixed. The air is such a mixture. Rock is usually a mixture of several minerals which are themselves true chemical compounds. Alloys are mixtures of different metallic elements; brass, for instance, being a mixture of the elements copper and zinc.

There are many thousands of compounds, but only about ninety elements. A table of elements is placed at the end of this book, and should be read through at this stage. Some elements are very common, such as oxygen and iron, while others are rare, like helium and thorium, and are found only in very small quantities. The numerous compounds may be grouped into classes such as oxides, sulphides, and so forth, depending on the nature of their composition. Such grouping makes a study of them a much simpler matter than it otherwise would be.

Chemical changes are usually represented by equations, or statements that name the substances started with and those resulting. The following word equations are a summary of the experiments and demonstrations thus far performed, and they should be studied, and the nature of the substances used and obtained should be recalled.

Copper and oxygen give copper oxide.
Iron and oxygen give iron oxide.
Zinc and oxygen give zinc oxide.
Tin and oxygen give tin oxide.
Magnesium and oxygen give magnesium oxide.
Sulphur and oxygen give sulphur dioxide.
Mercuric oxide gives mercury and oxygen.
Potassium chlorate gives potassium chloride and oxygen.
Carbon and oxygen give carbon dioxide.
Phosphorus and oxygen give phosphorus pentoxide.
Water gives hydrogen and oxygen.
Copper oxide and hydrogen give copper and water (hydrogen oxide).

DEFINITIONS

Decay is a process of slow oxidation dependent upon the action of certain kinds of bacteria in the air.

Rusting is a process of slow oxidation of iron dependent upon the action of moisture and carbon dioxide in the air.

Spontaneous combustion is a sudden and rapid oxidation caused by the accumulation of heat produced during a long period of slow oxidation.

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Reduction is the reverse of oxidation, and consists in removing oxygen from a substance.

Elements are substances that we cannot decompose into simpler substances by any means yet known.

Compounds are substances that we can decompose into simpler compounds, or into elements.

QUESTIONS

1. Under what class of changes does the charring of sugar in Experiment 1 belong?

2. What is the use of stove polish?

3. If you succeeded in changing a substance chemically by heating it, and were uncertain whether the change were one of oxidation or of reduction, how could you determine which it was?

4. Which substances studied are elements? which compounds?

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CHAPTER V

SIMPLE REPLACEMENT

In our effort to get information about the changes that matter is constantly undergoing we have discussed the action between elements which results in the formation of compounds, and also the actions by which compounds may be separated into their component parts or elements. We are now ready to study a third change which perhaps may seem quite complicated, and yet is a type that is very important in the understanding of the change discussed in the next chapter, the most common of all chemical changes.

If a little zinc be put into a test-tube and covered with dilute hydrochloric acid vigorous action immediately takes place. Little bubbles of gas seem to rise from the zinc through the liquid to the surface of it. Sometimes the zinc itself seems to rise, but a careful inspection shows that the bubbles cling to the zinc and so help lift the zinc through the liquid. The little bubbles are bubbles of gas. The rapid formation of gas in small bubbles in a liquid is called effervescence.

What is this gas formed when hydrochloric acid is put with zinc? Where did it come from? Putting a lighted splinter at the mouth of the test-tube produces a popping sound, or an explosive sound. Referring back to the experiment on the electrolysis of water the same phenomenon occurred when one of the gases into which water was decomposed was tested. That gas was hydrogen. The same conclusion may be drawn here — that hydrogen is the gas formed when hydrochloric acid and zinc are put together. Could this gas have come from the zinc? If we think about zinc we recall the fact that it is an element, and that an element is comparatively a very simple substance. We must conclude, therefore, that the gas could not have come from the zinc itself, but must have come from the only other substance involved, namely, the hydrochloric acid. Since the hydrogen differs from the acid, the hydrogen must have been set free from it by the action of the zinc.

If action between the acid and zinc continues long enough the zinc disappears completely. What has become of it? Another experiment will help solve this question. The contents of the test-tube should at this stage be filtered and allowed to evaporate completely. A white solid appears in the evaporating dish which is unlike either of the original substances. This solid is the compound zinc chloride. Its name suggests that the zinc is in this compound, and that the other part of the compound has some of the hydrochloric acid in it.

It is now possible to discuss the kind of action that takes place between the acid and the zinc. Zinc is an element, containing only zinc; therefore, no hydrogen could have been formed from it. The acid must have contained the hydrogen. Joined with that hydrogen to form the compound hydrochloric acid is the element chlorine. When the metal zinc reacts with the acid the resulting substances are hydrogen and zinc chloride.

Zinc plus hydrochloric acid gives hydrogen plus zinc chloride.

When such an action as this occurs it is called simple replacement; in other words, the zinc has crowded out the hydrogen from the acid, setting it free, and at the same time has slipped into its place making a new substance, zinc chloride, which is not seen at first because it is dissolved in water.

This kind of action occurs frequently and may take place between substances other than metals and acids. The next experiment involves some of these other substances, such as silver nitrate and lead nitrate compounds containing silver and lead respectively. Each of these compounds contains a metal, and a metallic element is allowed to interact with each. Again potassium iodide contains potassium, and iodine, a non-metal. Reacting with it is chlorine, a non-metallic element.

EXPERIMENT 5

Purpose.

To study simple replacement.

Directions.

a. Put three or four inches of magnesium ribbon in a test-tube. Add enough dilute hydrochloric acid to cover the magnesium.

Proceed in this experiment exactly as in that between zinc and hydrochloric acid.

Questions. What do you observe? Under what conditions have you observed a similar phenomenon?

NOTE. The second substance formed in this case is magnesium chloride.

b. Clean some copper wire by dipping it in hydrochloric acid and washing carefully in water. If not perfectly clean and bright use emery paper and wash again. Fill a test-tube one-third full of a solution of silver nitrate. Insert the copper in the solution of silver nitrate, being careful to leave a part of the metal untouched by the liquid. Let this stand until decided action has taken place (Fig. 10).

FIG. 10

Questions. What do you observe? Remove the copper; how does it compare in size with the copper in the beginning? What has the copper done to the silver?

c. Repeat *b*, using zinc and lead nitrate.

Questions. Same as for *b*, replacing words " copper " and " silver nitrate " with " zinc " and " lead nitrate."

d. Repeat *b*, using iron and copper sulphate.

Questions. Same as for *b*, replacing words " copper and silver nitrate " with " iron and copper sulphate." In *a, b, c, d* are the substances used elements or compounds? Discuss the kind of action in each case. Write the word equation for each.

e. To a few crystals of potassium iodide in a test-tube add water until the tube is about one-third full. Shake until all the potassium iodide has gone into solution, then add chlorine water. (This chlorine water should be prepared by the teacher before class time.) If necessary, add carbon bisulphide or chloroform — a few drops of either — in order to concentrate the color.

Questions. What is the color of the liquids before and after putting them together? What element is set free? Discuss the type of action. Write the word equation.

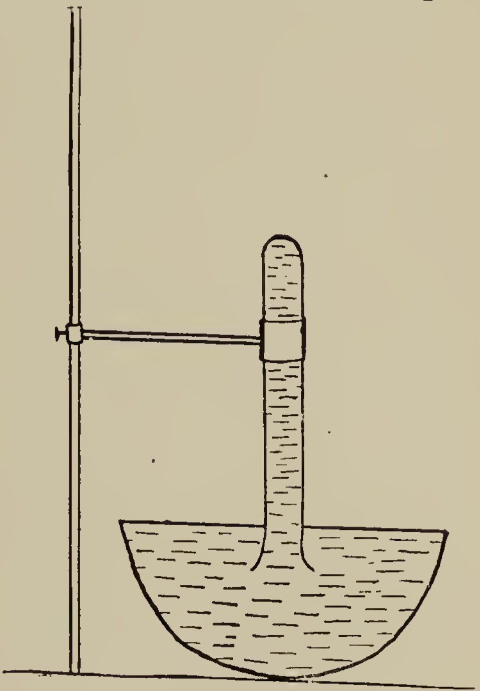


FIG. 11

(*To be done by the teacher*)

f. Invert a test-tube full of water in an evaporating dish of water, holding it in place by clamps (Fig. 11). Take a *small* piece of sodium, and thrust by means of tongs into the mouth of the test-tube. When all action has ceased remove the test-tube and its contents carefully, keeping the mouth of the test-tube covered with the thumb. Bring the tube still inverted near a colorless Bunsen burner flame, and remove the thumb.

Questions. What occurs? It is indicative of what substance? Where must this substance have come from? After evaporating to dryness the contents of the evaporating dish, is the residue like or unlike the original substance? It is called sodium hydroxide and contains sodium together with hydrogen and oxygen from the water. Write the word equation.

A careful study of all the reactions which took place in the last experiment shows that they are all of one type. In each case an element and a compound have been put together. In each case the resulting new substances have been an element and a compound, but the new element has been set free from the original compound, and the new compound always consists in part of the original elemental substance. Considering the original element only, it seems as if synthesis has taken place; considering the original compound only, decomposition seems to be the kind of action involved. When two substances react with each other to form two new substances it is not accurate to think of each substance by itself. Chemical actions probably do not take place in that manner. So to bring out more clearly the fact that these changes observed in the last experiment occur simultaneously to both the element and the compound the term "simple replacement" is given.

The various parts of the experiment show similarities and differences. Both magnesium and zinc act with hydrochloric acid; magnesium and zinc are both metallic elements. Some metals like iron, copper and zinc will replace respectively other metals like copper, silver and lead from their compounds. Why some metals act in this way and others do not may be more clearly understood by reading a later chapter. Again, a metal like sodium attacks water, forming hydrogen and sodium hydroxide. Only a few metals can do that, for, as is already known, water is a very stable compound. Only such agents as the electric current or a very active element like sodium can affect it. Potassium, another metallic element very similar to sodium, has the same ability; and also magnesium, iron and zinc, provided the temperature is high enough to help along the action. When the metallic elements react with water hydroxides are formed, and at the same time the part of hydrogen is set free. These hydroxides or bases possess definite characteristics, and contain chemically a metal, and part of the water with which the metal reacted. Thus we see that some metals act with water, some with acids, and some with metallic compounds.

In the experiment with chlorine and potassium iodide, the chlorine is a non-metallic element. It is gaseous in form, greenish yellow in color, has a very irritating and unpleasant odor, and since it is soluble in water is used in that condition for convenience. Both the chlorine and the iodine which it frees from the potassium iodide are members of what is called the halogen group, a group of four elements showing the same characteristics in gradually differing degrees.

Referring back to the demonstration between copper oxide and hydrogen given in Chapter IV, the element reacting is a non-metallic element. This may very properly be considered here also as an illustration of replacement. Considering the reaction as a whole, hydrogen has replaced the copper, so that while the copper oxide has been reduced, the type of action which produced this reduction is replacement. Both of these last two illustrations show clearly that non-metals as well as metals have replacing ability; and also that hydrogen either may be set free from water or acid, or may in its turn set free a metal. It will be interesting later to recall the fact that a metal in replacing another element always replaces a metal or hydrogen.

DEFINITIONS

Effervescence is the rapid formation of a gas within a liquid.

Simple replacement is the chemical action between an element and a compound in which the element takes the place of one of the elements in the compound.

Evaporation is the changing of a liquid to a gas. It is a process used frequently to separate a solid out of solution.

QUESTIONS

1. Compare the disappearance of salt in water with the disappearance of zinc in hydrochloric acid.
2. We sometimes say an acid " eats " a metal. What is really meant by this statement?
3. What is the kind of chemical action occurring when hydrogen " pops " as it is brought near a Bunsen burner flame?
4. In your own experience which seems more important, metals or non-metals? Why?
5. Give several reasons why magnesium will not make a good fuel.
6. Most metals are heavier than water. How can a steel warship float?
7. What are alloys? What is solder? Steel? Wrought iron? Cast iron?

CHAPTER VI

DOUBLE REPLACEMENT

EXPERIMENT 6

Purpose.

To study the preparation of a common acid — hydrochloric.

Directions.

a. Arrange the apparatus as in the figure, putting about five grams of sodium chloride in the flask, then inserting the stopper and filling the pan with water. Pour enough concentrated sulphuric acid through the thistle tube to cover the sodium chloride. Be sure that the end of the thistle tube is under the liquid. Heat the water bath to

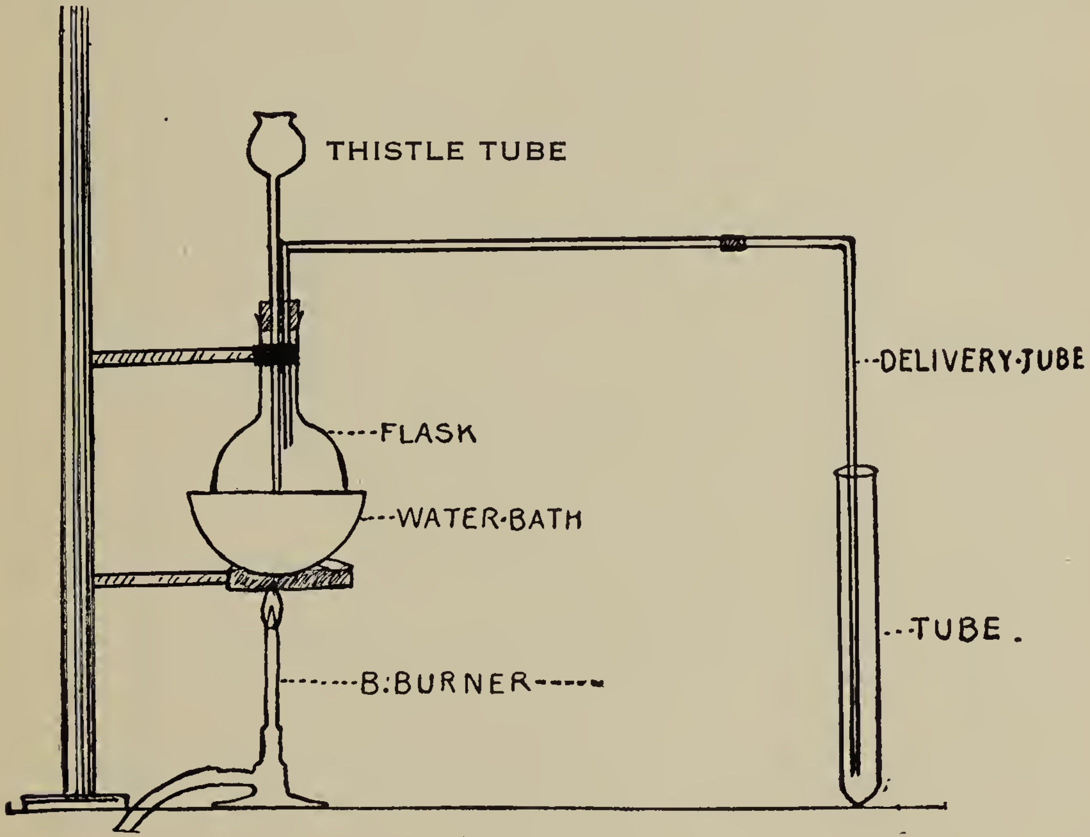


Fig. 12

boiling, noting any action which takes place in the flask. Moisten a piece of blue litmus paper and hold it at the end of the exit tube. When the color of the litmus paper changes, the acid is being formed in the flask. Then insert the delivery tube well into a dry test-tube. Let the gas pass into the test-tube for some time after a moist blue litmus paper held at its mouth changes color, to make sure that the tube is full of gas. Collect one test-tube of gas in this way, stopper it with a rubber stopper, and then pass the gas into a test-tube half full of water for five minutes.

b. Invert the test-tube of dry gas in cold water, remove the stopper, and leave in this position. Note what happens. When the action has stopped, slip the thumb over the mouth of the test-tube, remove from the water, and test the contents with litmus. Taste the contents by touching a glass rod to the solution and then to the tip of the tongue. Describe the taste.

c. Drop a piece of granulated zinc into the water solution of the gas obtained in part *a* by passing the gas into water for five minutes. Describe the action. Place a stopper very loosely in the test-tube, and after a few minutes remove it and place the mouth of the test-tube close to the Bunsen flame. Result?

Questions. Is hydrochloric acid gas soluble in water? What are the properties of the solution (taste, effect on litmus, action with zinc)? What part of this experiment proves that hydrochloric acid is a compound? Can you tell how this compound is formed, considering the nature of the original substances?

NOTE. These properties are the properties of many acids. Along with the acid is formed sodium sulphate. Write the word equation.

The preparation of hydrochloric acid illustrates the fourth type of chemical action, double replacement. Both of the substances used, sodium chloride and sulphuric acid, are compounds. When brought together under the influence of heat a change takes place which results in the formation of hydrochloric acid. At the same time that the acid is formed sodium sulphate is also formed in the flask.

Sulphuric acid plus sodium chloride gives hydrochloric acid plus sodium sulphate.

Thus there are two compounds formed by this change, compounds which are different from the original compounds and yet which must necessarily contain parts of them. The word equation, which represents what happened in the flask, shows that a part of the compound sodium chloride, the chlorine, went to the forming of the hydrochloric acid, and a part of the sulphuric acid went to the forming of sodium sulphate. In simple replacement one element crowds out another from a compound; in double replacement both compounds are split up, and the separated parts reunite in another way. This action is more clearly comprehended if the two compounds are thought of as separated into component parts first, followed by the recombination of these parts as a second stage in the process. In reality the reaction takes place as a whole.

Double replacement is concerned always with compounds. Since compounds may be complex substances, double replacement may often be a very complicated change. In the laboratory, changes between two compounds are studied because such a study is for illustrative purposes. In the large changes in nature more than two compounds will be involved. Such a change as takes place when the contents of the white and blue papers of a Seidlitz powder are put together makes use of two compounds only, but a cake is produced because action has taken place among several compounds. Again, when food is taken into the digestive system, and the various juices act upon it, changes must take place simultaneously among several compounds in order to get the food ready for assimilation into the blood supply. The chemistry of foods furnishes many illustrations of this sort.

Although double replacement can only take place between compounds, not every compound will act with every other compound. After the types of reactions have been studied, and when in addition to that all the kinds of compounds have become familiar, it will be possible to predict pretty well when action will take place, and what substances will be formed, without actually trying them in the laboratory. In the laboratory we find out how general classes of substances react; later we shall endeavor to explain why these changes and reactions occur. Before explanations can ever be given we have to know that with similar causes similar results follow. That is what the laboratory work is trying to do, and when we have all those facts we shall be ready for explanations.

Laboratory work not only makes it possible to know about the kinds of reactions, but also makes it possible to know the characteristics of the substances themselves. In the making of hydrochloric acid the type of reaction was important because it represented the first example of double replacement. From another point of view the acid itself is an important factor. Acids are very common substances, and so important that it has been said that the civilization of a country depends upon the amount of sulphuric acid which it produces. Many of the acids are very necessary for commercial purposes, like hydrochloric, others are important as constituents of our food, like the citric acid of lemons. Most people think of acids as liquids and as substances that are corrosive. Acids, when pure, may be liquid like sulphuric, they may be gaseous like hydrochloric, and sometimes they are solid like citric. They may be corrosive like nitric or soothing like boracic, which is often used as an eye wash.

When hydrochloric acid was poured upon zinc, effervescence ensued which meant that a gas was being formed, the gas hydrogen. Hydrogen is a constituent of hydrochloric acid, and is set free from it by the zinc. Other acids, like sulphuric, treated in the same manner, give the same result. So if many of the common acids be tested hydrogen is found to be present in each one of them. The hydrogen may not always be set free by a metal, but by one experiment or another hydrogen is found to be an essential part of all acids. Since hydrogen is an element common to all acids, there may be other common likenesses. One of these is the taste. If a drop of acid be put on the tongue a characteristic sour taste results, the sour taste differing in degree from that given by the acid in oranges to the sourness of hydrochloric acid. In fact, the terms "sour" and "acid" are often in the everyday world used synonymously.

If solutions of these acids be tested with litmus paper all of them will give the change from blue to pink, differing in degree with the acid used. Litmus is a substance made from a lichen which if put into solution with alcohol has the power to turn pink when a drop of acid is added. This is a quick way of detecting the presence of an acid, but it does not tell what acid it may be.

Chemically all acids are compounds, having hydrogen as one of the constituent elements. Besides hydrogen all acids contain either a non-metallic element, giving what is called a binary acid like hydrochloric acid, or a group consisting of a non-metallic element and oxygen, making what is called a ternary acid like sulphuric acid. This combination of a non-metallic element with oxygen acts as a group in all reactions involving acids, and for that reason is spoken of as a radical. The radical of an acid thus distinguishes the acid.

That an acid must contain a non-metallic element can be shown by preparing sulphur dioxide in a bottle just as it has been prepared before, then adding water to the gas. Sulphur dioxide dissolves in water and when tested with litmus paper gives the characteristic pink color. The acid formed is sulphurous acid. In a like manner phosphorus pentoxide obtained in Experiment 4, when phosphorus was burned in oxygen dissolved in water, will give the litmus test for acids. In this case phosphoric acid is formed. These two examples are sufficient to show that from a non-metallic element one may under the proper conditions build up an acid. The word equations are:

Sulphur plus oxygen gives sulphur dioxide.
Sulphur dioxide plus water gives sulphurous acid.
Phosphorus plus oxygen gives phosphorus pentoxide.
Phosphorus pentoxide plus water gives phosphoric acid.

The next question to settle is whether metallic elements are acid formers, or whether their oxides with water give a different class of substances.

The experiments given in Chapter V, when sodium and potassium act with water, showed that besides hydrogen another substance was formed called a base, or hydroxide. If also magnesium oxide, formed as in one of the first experiments, be dissolved in water the same sort of a substance is produced. The word equation is:

Magnesium plus oxygen gives magnesium oxide.
Magnesium oxide plus water gives magnesium hydroxide.

To prove quickly the presence of these substances we use litmus paper. But in all these cases the pink litmus is turned blue, indicating the presence of a substance whose effect on litmus is just opposite to that of an acid.

These hydroxides or bases are all formed in such a way that the metal is known to be an essential part, and in all cases water enters into the reaction. The chemical experience of many people proves definitely that all bases contain a metal, together with the elements oxygen and hydrogen. The presence of oxygen and hydrogen in all bases gives them their common characteristics, such as their effect on pink litmus, the bitter taste, the slippery feeling to the touch, and their ability to dissolve fats, just as the presence of hydrogen in all acids gives them their characteristics. Like acids some bases are much more common than others, sodium hydroxide, potassium hydroxide, calcium hydroxide and ammonium hydroxide being the most common. The name hydroxide is the chemical term for base. These bases are all soluble, and find wide use in our everyday life. Bases, then, are compounds containing a metal, hydrogen and oxygen. The hydrogen and oxygen stay together in a group when a base acts chemically, and so like the acid group is called a radical. In the case of acids the radical distinguishes one from the other, the hydrogen giving the common characteristics; in the case of bases the metal makes the distinction, while the basic or hydroxyl radical gives the common characteristics.

Since the four bases mentioned in the last paragraph are all soluble it is comparatively easy to prepare them. In the experiment with sodium and water after the hydrogen was collected the solution was evaporated to dryness, giving a white residue, sodium hydroxide. Potassium hydroxide may be prepared in a similar fashion, and calcium hydroxide by adding calcium oxide to water. The next experiment gives a method for making an insoluble base, not so much for the importance of the substance as for the getting of much general chemical information.

The terms metal and non-metal have been used so far without any exact definition. Metals are of such common occurrence and their physical properties of hardness, luster, malleability and ductility are so well known that they do not need emphasis. Not all metals, however, possess all these qualities. Sodium, for instance, is soft, not hard; mercury is a liquid metal, not a solid. These physical characteristics of luster, hardness and so on, are not always a means for distinguishing the metals from the non-metals. Iodine has a well-defined metallic luster, but the chemist classes it with the non-metals. A chemical test then is better. It has already been said that the oxides of metals produce bases when added to water, and that the oxides of the non-metals with water produce acids. Acids and bases are such well defined classes of substances that this means of distinguishing a metal from a non-metal is much more uniform, and much to be preferred.

EXPERIMENT 7

Purpose.

To study the preparation of insoluble bases.

Directions.

Half fill a test-tube with a solution of aluminum sulphate, then add a solution of the base made in the previous experiment, sodium hydroxide. Result? This insoluble solid is aluminum hydroxide.

Fold a filter paper in fourths, fit it into a funnel, wet, and press gently against the funnel until it fits as smoothly as possible. Pour the contents of the tube into the paper, and catch what passes through in another test-tube.

Repeat with iron sulphate and sodium hydroxide.

Questions. What did you observe to form in both tubes? What was alike in the two cases, and what different? What is the condition of the substance going through the filter paper? If this change illustrates double decomposition, what else is formed besides the hydroxides? Write the word equation.

NOTE. The formation of such insoluble substances is called the formation of a precipitate. The liquid passing through the filter paper, often containing solids in solution, is called the filtrate.

The study of double replacement completes the study of the kinds of changes which matter may undergo. Changes which are so familiar to us in the outside world are not so simple as the ones studied in the laboratory; in most instances they are combinations of two or more types, and for that reason make the study of them too difficult for a beginner. In the study of the four types of reactions several classes of substances have been brought to our attention: metals, and non-metals into which the elements may be grouped; oxides, acids, and bases into which compounds may be grouped. Each of these classes has its characteristic relation to the others. There remains only one more class for study, but it is so large a class and so important that it needs a chapter by itself for proper discussion.

DEFINITIONS

Double replacement, or double decomposition, is chemical action between two compounds which results in the formation of two new compounds by an exchange of elements.

An **acid** is a compound containing hydrogen that can be replaced by a metal. It is sour in taste, and turns blue litmus paper pink.

A **base** is a compound made up of a metal, hydrogen and oxygen. The hydrogen and oxygen are always in the same proportion. It turns pink litmus paper blue.

The terms hydroxide and base are synonymous.

A **metal** is an element which either replaces hydrogen or may be replaced by it, and whose oxide may be regarded as forming bases with water.

A **precipitate** is an insoluble solid formed when certain solutions react chemically. Most precipitates settle quickly to the bottom of the containing vessel.

Filtration is the process by which a solid mixed with a liquid may be separated from the liquid.

A **filtrate** is the clear liquid which is left as a result of filtration. It may or may not contain a substance in solution.

QUESTIONS

1. Why do green apples taste sour?
2. Sometimes soil is said to be sour. What does this mean? How could a particular soil be tested to determine whether it is sour or not?
3. Acids " burn " holes in one's clothing. What connection has it with ordinary burning?
4. Some acids burn the skin. What does this mean?
5. Under what conditions does milk sour? What kind of a substance is formed? Can you prove it?
6. Is the earthly material which separates out of muddy water when it stands a precipitate? Give reason for your answer.
7. What is the foundation material of agate ware? If fruits are cooked in an agate dish from which the agate has worn off what chemical action would take place?

CHAPTER VII

SALTS

EXPERIMENT 8

Purpose.

To study neutralization.

Directions.

Measure out 25 c. c. of a solution of sodium hydroxide into an evaporating dish. Add 25 c. c. of dilute hydrochloric acid. Note the temperature of the dish. Label two test-tubes " acid " and " base " respectively, and half fill each with its proper solution. Use a separate glass tube for each to drop the solution from. Stir the contents of the evaporating dish well, and test with blue and pink litmus. Then add a drop of the needed solution, stir well, and test again, adding a drop at a time, stirring and testing until neither litmus is changed. The solution is then said to be neutral. Put the evaporating dish with its contents over the Bunsen burner flame and evaporate *gently* to dryness. If any sediment or color appears, filter before evaporating. Some members of the class should pour their solutions into crystallizing dishes, label, and set aside to evaporate slowly.

Taste a bit of the residue. Do you recognize it? Compare this end product with the original substances. Use residue from both quickly and slowly evaporated solutions.

Questions. If this is salt, sodium chloride, what is the other substance formed? What type of reaction is this? Write the word equation.

NOTE. The acid and base solutions should be made by the teacher, about forty grams of solute to a liter of water.

The action of a base on an acid results in the formation of water and a salt. In the experiment the salt formed was common table salt, a very familiar substance. From the strictly chemical point of view perhaps the water is the more important substance; from the everyday point of view the salt is more important.

A salt is usually white in color, crystalline in form, and normally does not react with either kind of litmus paper. Chemically any salt is composed of the metal of a base, and all of an acid except the hydrogen. Since a salt has in it a part of a base and a part of an acid it is easy to understand why a salt does not act upon litmus, or in other words is a neutral substance. When acids and bases were studied some acids and some bases were found to be more active than others. If a base and acid of nearly equal activity are put together the resulting salt is neutral. If, however, an active acid and a less active base are put together the resulting salt is acid to litmus. When an active base and a less active acid are put together the resulting salt is basic (or alkaline) to litmus. Such salts are called acid or alkaline with respect to their action on litmus.

1. Quartz crystals. 2. Copper sulphate crystals. 3. Sodium chloride crystal (halite). 4. Calcium carbonate crystal (calcite). 5. Alum crystals.

The salts formed by the action of an acid on a base, or by the process of neutralization as it is called, are all soluble in the water which is formed at the same time and also any water which may have been used to dissolve the acid or the base. Because they dissolve in water it is possible to separate the solid salt in a very pure form in the shape of crystals by allowing the water to evaporate. These salts vary in their ability to dissolve, some dissolving in large quantities and some in small quantities in any given amount of water at any given temperature. When a given amount of water has in it all of the salt that it can hold at the given temperature, while some undissolved salt is still present, the solution is said to be saturated. Sometimes a saturated solution allowed to cool undisturbed will remain a solution. If, however, a small crystal of the salt be dropped in it, the salt crystallizes out immediately. Just before crystallization takes place the solution is said to be supersaturated. The best crystals are formed by using a saturated solution and allowing it to cool slowly. The more slowly the solution cools and the more slowly evaporation takes place the more perfect the crystal. Each salt has its own kind of crystal, just as the salt crystals formed in the experiment are cubical in form. A crystal, then, is often a transparent substance occurring in definite regular shape, and the process of getting crystals from solution, or from a molten substance, is called crystallization. Sometimes water stays in the crystal as it forms. Such water is not apparent to the naked eye, but makes itself known to the chemist when conditions are right. Such a crystal is said to contain water of crystallization. In some cases rapid evaporation of the solvent makes a crystal too small to recognize with the naked eye. In other cases the solid separates out in an amorphous state, that is, non-crystalline. Whenever the water of crystallization is driven out of a salt the result is an amorphous condition of that salt. It is then said to be anhydrous, or without water.

Neutralization is the simplest way of making a salt, but for many reasons it is not the way by which all salts are made. Often other methods are more convenient, either because other substances are more easily procured or because salts are formed as a by-product in the manufacture of other substances. Looking back to the experiment showing the action of a metal on an acid in Chapter V, or to the one showing the preparation of an insoluble base in Chapter VI, we find that along with the substances under discussion a salt was also made. In both of these illustrations the salt was for us in our study a by-product. The next two experiments show other common methods for making salts.

EXPERIMENT 9

Purpose.

To study the preparation of a salt.

Directions.

Put a few pieces of copper oxide (black granular form) into a test-tube one third full of dilute hydrochloric acid, or enough acid to dissolve the oxide. Shake and heat the contents. Filter if necessary, and evaporate very gently to dryness. What color is the residue? If dark in color, allow it to cool slowly. What happens? Can you account for it? Some students should evaporate slowly to obtain crystals.

Questions. What is the name of the residue left after evaporation? What is the type of reaction? Write the word equation.

EXPERIMENT 10

Purpose.

To study the preparation of a salt.

Directions.

Fill a test-tube about one-third full of a solution of barium chloride. Add silver nitrate solution slowly until no further action takes place. Filter and let the filter paper with its contents stand in the light for several hours. Evaporate the filtrate slowly to dryness.

Questions. What happens when these two salts are first put together? What does such a phenomenon indicate? Did the filtrate contain anything in solution? What is the type of reaction? Write the word equation.

NOTE. In this reaction there are two new salts formed.

Experiment 9 shows that an oxide may be dissolved in an acid thereby giving a salt and water. Since water is the only substance formed besides the salt it is an easy method to use. In this particular experiment the oxide of copper is much more easily procured than the hydroxide, and so as with many other metals this method of making its salt is preferable.

In Experiment 10 two soluble salts are put together, giving two new salts. That method has advantages and disadvantages. Two soluble salts only can be put together, and then to procure either one of the new salts in a pure state there must be some good way of separating them. In the experiment one was insoluble, therefore filtration separates them easily. Sometimes both salts formed are soluble in different degrees so that by continued crystallizing one may be obtained in a fairly pure condition. If on the other hand two salts were made by this method whose solubilities were close together, neither one could be made in a satisfactory manner. Because this is true there are other methods for making salts which are given in the summary below. Many of these methods are the ones actually used for the making of the salt, because the salt is the by-product in the manufacture of certain commercial products.

Below is given a brief summary of the methods.

1. Action between two elements, like sulphur and iron.
2. Dissolving an oxide in an acid, like copper oxide and hydrochloric acid.

3. Action of a base on an acid, like sodium hydroxide and hydrochloric acid.

4. Action of an acid on a metal, like zinc and hydrochloric acid.

5. Action of an acid on a salt, like sulphuric acid on sodium chloride.

6. Action between two salts, like barium chloride and silver nitrate.

7. Action between a metal and a salt, like copper and silver nitrate.

Which of the above methods may be used in the preparation of any given salt depends upon several facts. The solubility of the salt to be made has to be considered together with the occurrence of the substances which react to form that salt, and also the state of purity desired.

So far all the compounds studied belong to the classes of substances called oxides, acids, bases, or salts. Of all these the salts are by far the greatest in number. A very large proportion of the substances we come in contact with in our daily lives belong to the class of salts. The mineral substances in our food, the substances which make water hard, the material of our bones and teeth, many of the medicines used by physicians, the substances making up the largest part of the earth's hard outside crust, are all salts. Everywhere we turn the substances we so commonly find and so frequently use are salts. Several salts have come to our attention — we know even a new compound to be a salt because it contains a metal and an acid radical. Sodium chloride, one of the earliest salts known and from which the group name came, barium nitrate, silver nitrate, lead nitrate and potassium bromide, which have been used in previous experiments, are now known to be salts. And in a later chapter will be given many other common substances which chemically are salts.

In the early beginnings of scientific knowledge the chemical nature of acids, bases and salts was not so thoroughly understood as it is at the present time. Substances which now seem so simple were then considered not so simple apparently, and names were given to them in a very arbitrary fashion. As the number of known substances increased, and more was known of the chemical nature of these substances, the original names were often very confusing. For instance, potassium sulphate had at one time five different names. Lavoisier is largely responsible for beginning the system of nomenclature which holds at the present day. By this system one is able to name a compound when knowledge of its chemical relationship becomes familiar. The most common acids are hydrochloric, hydrosulphuric, sulphurous, sulphuric, nitrous, nitric, carbonic, citric, oxalic and tartaric. If an acid has in it two elements only, the name of the acid begins with the prefix " hydro " and ends in " ic," as hydrochloric acid. The salt of that acid has for its last name " chloride," the word ending in " ide." If the acid has oxygen in it in addition to the hydrogen and the non-metal, its name depends upon the amount of oxygen in it. Sometimes there are two acids containing the same elements but in different proportions. In such a case the acid containing the smaller amount of oxygen has its name ending in " ous," like sulphurous or nitrous. The salts of such acids have the last name ending in " ite," as sulphite or nitrite. The acid containing the larger proportion of oxygen has its name ending in " ic," as sulphuric acid or nitric acid. Salts of such acids bear the name ending in " ate," such as sulphate or nitrate. In the case of salts the first part of the name is the name of the metal of the base to which it is related.

Just as acids vary in the amount of oxygen they contain, so oxides of certain metals may vary in the amount of oxygen they contain. Metals like copper, iron and lead form two oxides. When such is the case the endings "ous" and "ic" are again used to show their chemical composition. This is shown by the two oxides of copper, cuprous oxide and cupric oxide.

SUMMARY

Acids may be solids, liquids or gases. They are for the most part soluble, turn blue litmus pink, taste sour, and contain hydrogen replaceable by a metal. Besides hydrogen, acids always contain non-metals.

Bases are mostly solids, only a few are soluble. They turn litmus blue, have a bitter taste, slippery feeling, and contain a metal, hydrogen and oxygen. The hydrogen and oxygen are always present atom for atom.

Salts. Salts are usually white crystalline solids (except iron and copper). They are usually soluble, and neutral to litmus. They contain a metal and an acid radical.

EXPERIMENT 11

Purpose.

To determine the effect of some common substances upon litmus.

Directions.

a. Make solutions of the following: cooking soda, washing soda, cream of tartar, borax, soap, lime, cleaning powder, molasses. Then test each with red and blue litmus and record result.

b. Test the following without making solutions: sour milk, tooth paste, saliva.

Record results.

DEFINITIONS

A **radical** is a group of elements forming part of a compound, which remain together throughout many chemical changes.

Neutralization is the term applied to the action of a base on an acid whereby a salt and water are formed.

A **salt** is a compound composed of a metallic element and an acid radical.

A **crystal** is a solid substance bounded by plane surfaces of regular shape and arrangement.

An **amorphous** substance is any substance which is non-crystalline.

Crystallization is the process by which a substance is formed in crystals. It is usually brought about by dissolving the solid in some liquid and allowing the liquid to evaporate.

A **saturated solution** is one which does not change in composition when in contact with some of the solid.

A **supersaturated solution** is one which is more concentrated than a saturated solution.

A **solvent** is any substance which will hold other substances in solution.

A **solute** is any substance which can be made to dissolve.

QUESTIONS

1. If not enough silver nitrate is added to barium chloride what substances will be in the filtrate?

2. Potassium and sodium nitrates are soluble salts. How can you account for their being found in the earth? Will the same explanation hold for sodium and potassium chlorides?

3. Is the action between an oxide and an acid double replacement?

4. Why cannot a salt be made from phosphorous pentoxide and an acid? Why not from copper oxide and a base?

5. Does sugar on grape fruit neutralize the acid or merely cover up the sour taste?

6. What name is given to the salts of hydrobromic acid? (Hydrobromic is very similar to hydrochloric acid.) To the salts of phosphoric acid? To the salts of phosphorous acid? To salts of chloric acid?

7. Do acids affect the teeth?

8. Refer to experiment on testing common substances. How does the tooth paste result agree with No. 7?

9. Why is baking soda sometimes used for a "sour stomach?"

10. Why is putting lime on some soil said to sweeten it?

11. Why is baking soda put with molasses to make gingerbread and molasses candy?

12. In the experiment on testing common substances which are used as cleansing agents? What common property have they?

13. Why is baking soda used with sour milk and not with sweet milk in cooking?

CHAPTER VIII

INTRODUCTION TO THEORY

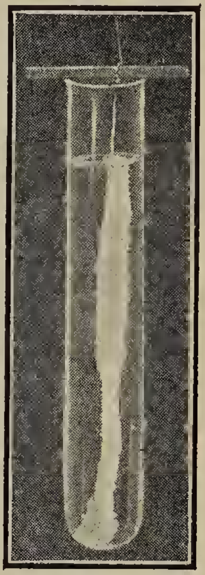
WE shall now resume the study of chemical change, which is the foundation of a knowledge of chemistry. An acquaintance with common chemical substances is necessary to a knowledge of household chemistry, but to emphasize it as the major requirement would show more interest in effects than in causes. Let us then recall the experiments which produced some of these substances, and review rapidly the signs which told us that new substances were forming.

Zinc dust and powdered sulphur, mixed thoroughly and ignited. A chemical action producing much light yet not an oxidation.

Zinc and acid effervescing.

Probably the most common sign of a chemical change is the production of heat. Every case of burning is a pronounced illustration of this. Sometimes light accompanies heat and gives additional evidence of change, as when a match is struck or a gas jet lighted. However, if the action of sodium on water, demonstrated among the experiments in replacement, be recalled, and the neutralization of sodium hydroxide by hydrochloric acid, it will be seen that heat alone is often sufficient to determine that a chemical change is taking place. Efferverscence, or the formation of a gas, is often evidence of a chemical change, as when hydrochloric acid was poured on zinc. Another sign is the formation of a solid precipitate when two solutions are poured together. This may well be seen when a solution of mercuric chloride is poured into a solution of potassium iodide. A new substance is formed which is insoluble and so separates out of solution. Similar cases were the replacement of silver in silver nitrate by copper, and the other experiments under replacement in Chapter V. Finally, mere change of color or change of odor may denote a chemical change, as when leaves turn in the autumn or food decays.

a and *b*. Solutions of barium chloride and silver nitrate. *c*. Precipitate of silver chloride formed on mixing *a* and *b*.



Silver tree. Crystals of silver forming on a copper wire suspended in a solution of silver nitrate.

All these, however, are merely outward signs by which it is sometimes known that new substances are being formed. There is no guide, so far, to tell us what substances will react with each other and what will not. A few simple experiments will make this plain.

EXPERIMENT 12

Purpose.

To study reactions and no reactions.

Directions.

Take half a test-tube of dilute solution of each of the following compounds and a small piece of each of the metals. Place the substances together as directed, describe what happens, looking for any signs of chemical change. If nothing occurs make a note of the fact. Where a reaction takes place write the word equation.

Compounds: Sodium hydroxide, hydrochloric acid, ferric nitrate, potassium nitrate, potassium chloride, copper sulphate, ferrous sulphate, calcium carbonate, calcium sulphate. Use the last two dry.

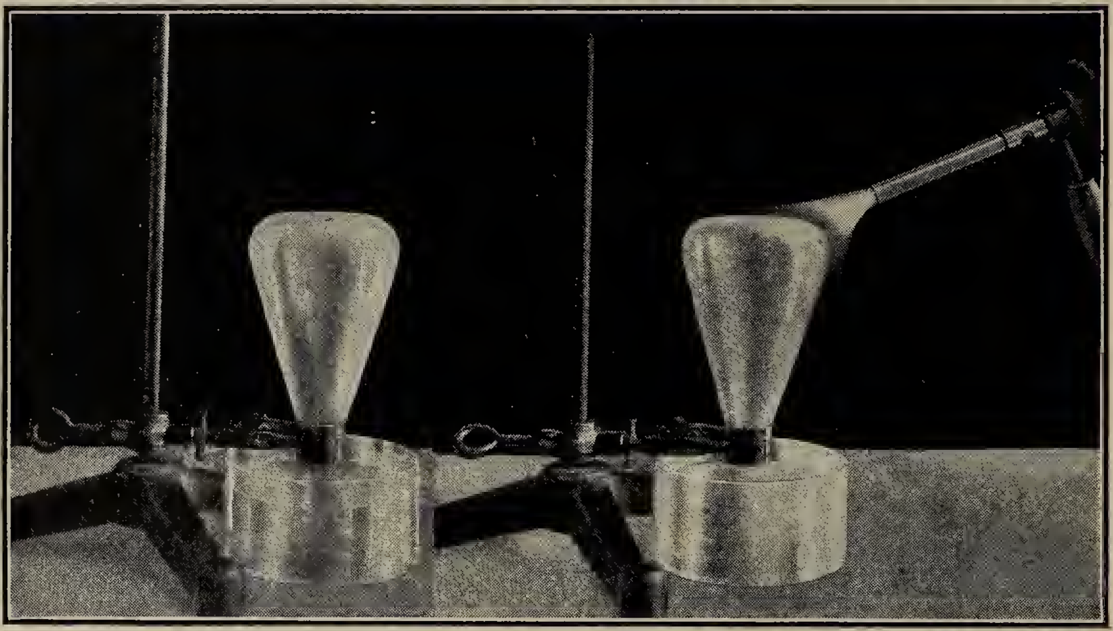
Metals: Copper, zinc, iron, cleaned by sand-papering.

Place together

a. Sodium hydroxide and ferric nitrate.
b. Sodium hydroxide and potassium nitrate.
c. Hydrochloric acid and zinc.
d. Hydrochloric acid and copper.
    Allow *c* and *d* to stand for five minutes.
e. Sodium hydroxide and hydrochloric acid (hydrogen chloride).
f. Sodium hydroxide and potassium chloride.
    Feel the temperature of *e* and *f*.
g. Iron and copper sulphate.
h. Copper and iron sulphate.
    Allow *g* and *h* to stand for at least five minutes.
i. Hydrochloric acid and calcium carbonate (marble).
j. Hydrochloric acid and calcium sulphate (gypsum).

Questions. Compare *a* and *b*, *c* and *d*, *e* and *f*, *g* and *h*, *i* and *j*. In which did reactions take place, in which not? Can you see any reason for the difference?

To have explanations of these phenomena, to know when substances will interact and when not, a study is necessary of the theories which have been devised by physicists and chemists in order to explain them; and if any of these theories seem hard to believe, it must be remembered that they not only explain these and many other happenings, but have proved useful hypotheses in the light of which new facts have been discovered and new substances found. Theories need to be modified and revised from time to time as new discoveries demand, but the most incredible theory has sometimes received startling confirmation at the hands of the experimenter.



Heat doing Mechanical Work

In the first picture, air pressure on the water in the dish is holding the water up in the flask. In the second picture, the heat from the flame has driven the water down.

The study of theory must be prefaced by a few statements concerning energy. Here there is only enough time to state that energy is the ability to do work; that there are various ways in which energy may be shown, such as mechanical energy, light, heat, electrical energy, chemical energy; and that these may be changed into one another and back again, but that the total amount of energy in a given experiment or combination of circumstances may be neither diminished nor increased. This fact is called the law of the conservation of energy, and is comparable to the law of the conservation of matter mentioned earlier. Heat is the commonest form of energy, and the form into which other forms are most easily converted. Indeed it is difficult, in trying to obtain any other form of energy, not to lose some of it as heat. Illuminating gas and electric currents are instances of this. It is necessary in experimental work to measure the heat evolved or used, and for this purpose a unit of heat called a calorie has been agreed upon. A calorie is that amount of heat which will raise one gram of water one degree Centigrade in temperature. A Calorie is equal to one thousand calories.

Light doing Work

Blueprints of sweet gum and sycamore maple leaves, showing the work done by sunlight in producing a chemical change in the paper where it was unprotected by the leaf.

In chemical reactions heat is either given out by the substances in action or taken in by them. Most of the common reactions are of the former kind, in which heat is given out, and are called exothermic. Common burning or oxidation, for instance, is exothermic, and the fact that heat must usually be applied to the substance before it will burn should not confuse the matter. The substance cannot unite with oxygen below a certain temperature called its kindling point, but after the action starts the heat evolved is itself sufficient to keep up the temperature and permit the action to go on. Spontaneous combustion, it will be recalled, occurs when the heat evolved by slow oxidation of the oil on paint-stained cloths has accumulated in the pile until it is hot enough to ignite the cloths or other refuse. The amount of heat given out by the formation of a fixed weight of a substance, such as an oxide, is called its heat of formation, and is always the same for that substance. When heat is absorbed in a chemical reaction the reaction is called endothermic, and if a compound is formed it is said to have a negative heat of formation. Acetylene gas and ozone are examples of substances having negative heats of formation. As would be expected such substances are unstable, and readily break up, evolving the heat which they absorbed when made. Endothermic reactions, however, are commonly those of decomposition, resulting in simpler compounds or elements. The decomposition of water into hydrogen and oxygen is an example of endothermic reaction absorbing a large amount of energy.

Heat of formation is an important subject, for it has been found that in any chemical reaction where no external heat is applied, that substance will tend to form that has the greatest heat of formation. This principle explains the replacement of iodine in potassium iodide by chlorine in Experiment 5 *e*. The heat of formation of potassium chloride is greater than the heat of formation of potassium iodide, and when by mixing chlorine water and a solution of potassium iodide there is an opportunity given for its formation, the iodine is displaced by the chlorine. A mixture of iodine and potassium chloride solution would show no reaction.

The subject of heat had much to do with the development of the first theory we shall discuss, for it was due to observations of the effect of heat and pressure on gases that the theory was developed.

DEFINITIONS

A **scientific theory** is an explanation in general terms of the reason underlying phenomena.

Energy is the ability to do work.

The **law of the conservation of energy** states that we can neither create nor destroy energy.

A **calorie** is that amount of heat necessary to raise one gram of water one degree Centigrade. A **Calorie** is equal to one thousand calories.

Exothermic reactions are those in which heat is evolved.

Endothermic reactions are those in which heat is absorbed.

The **heat of formation of a substance** is the number of calories evolved when a fixed weight of the substance (the gram molecular weight) is formed. The heat of decomposition is the reverse of this.

QUESTIONS

1. Review the experiments already performed and tabulate them under the following heads: Those marked by (*a*) heat, (*b*) light, (*c*) change of color, (*d*) effervescence, (*e*) precipitate, (*f*) any other signs.
2. Is the effervescence of a charged water evidence of a chemical change or not?
3. Name as many different forms of energy as you can, and give an illustration of work done by each.
4. Give instances in which other forms of energy are converted into heat.

5. Give illustrations of exothermic and endothermic reactions other than those mentioned in the text.

6. Which kind of reaction, exothermic or endothermic, do you think that synthesis usually is? Decomposition?

7. In Experiment 2 in Chapter II, which contained more energy, copper and oxygen separately or the copper oxide formed from them? Is this an exothermic or an endothermic reaction?

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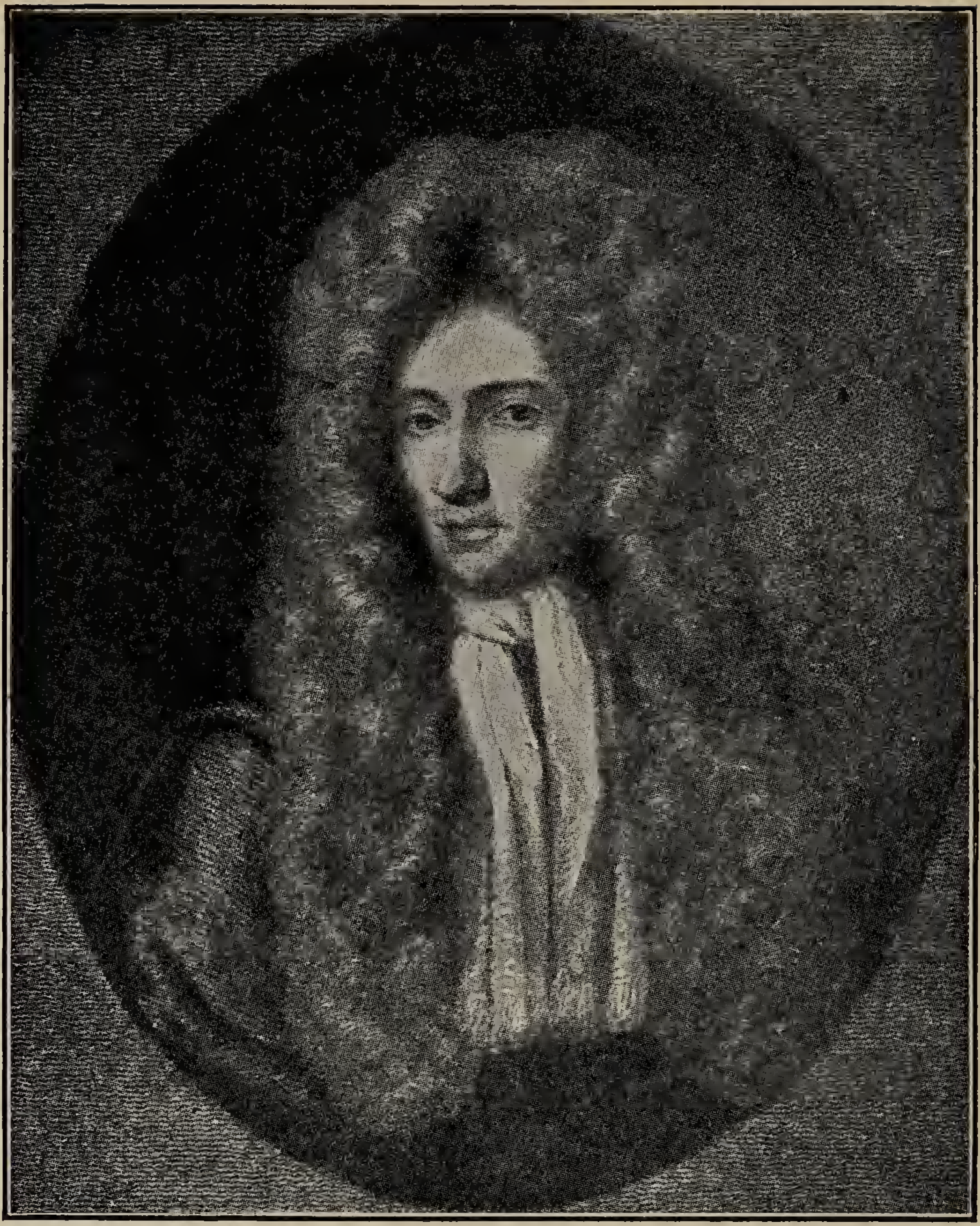
CHAPTER IX

MOLECULES AND ATOMS

It was found by repeated experiments that all gases of whatever composition behave alike when subjected to like temperature and pressure, expanding the same proportions of their volumes when equally heated, and contracting to equal fractions of their volumes when subjected to the same pressure. This suggested the idea that all gases were probably made up of very small particles, invisible even under the microscope. The particles of each kind of gas were thought to be alike in every way, including chemical composition and weight, were thought to have all the properties of the gas, and to be the smallest particles of it that could exist independently. These particles were supposed to be in constant motion or vibration, and constantly colliding with the sides of the containing vessel. This motion was thought to be readily changed into heat. It is evident that this explains the expansion of gases when heated, for the heat would increase the motion of the particles, and they would thus strike more violently against the walls of the container and so separate further from each other, increasing the volume of the gas. It explains also the rise in temperature when the gas is subjected to increased pressure, for forcing the particles closer together means that their motion would be restricted and that some of it would show itself in increased temperature. The theory was found later to apply equally well to liquids and solids as to gases, the



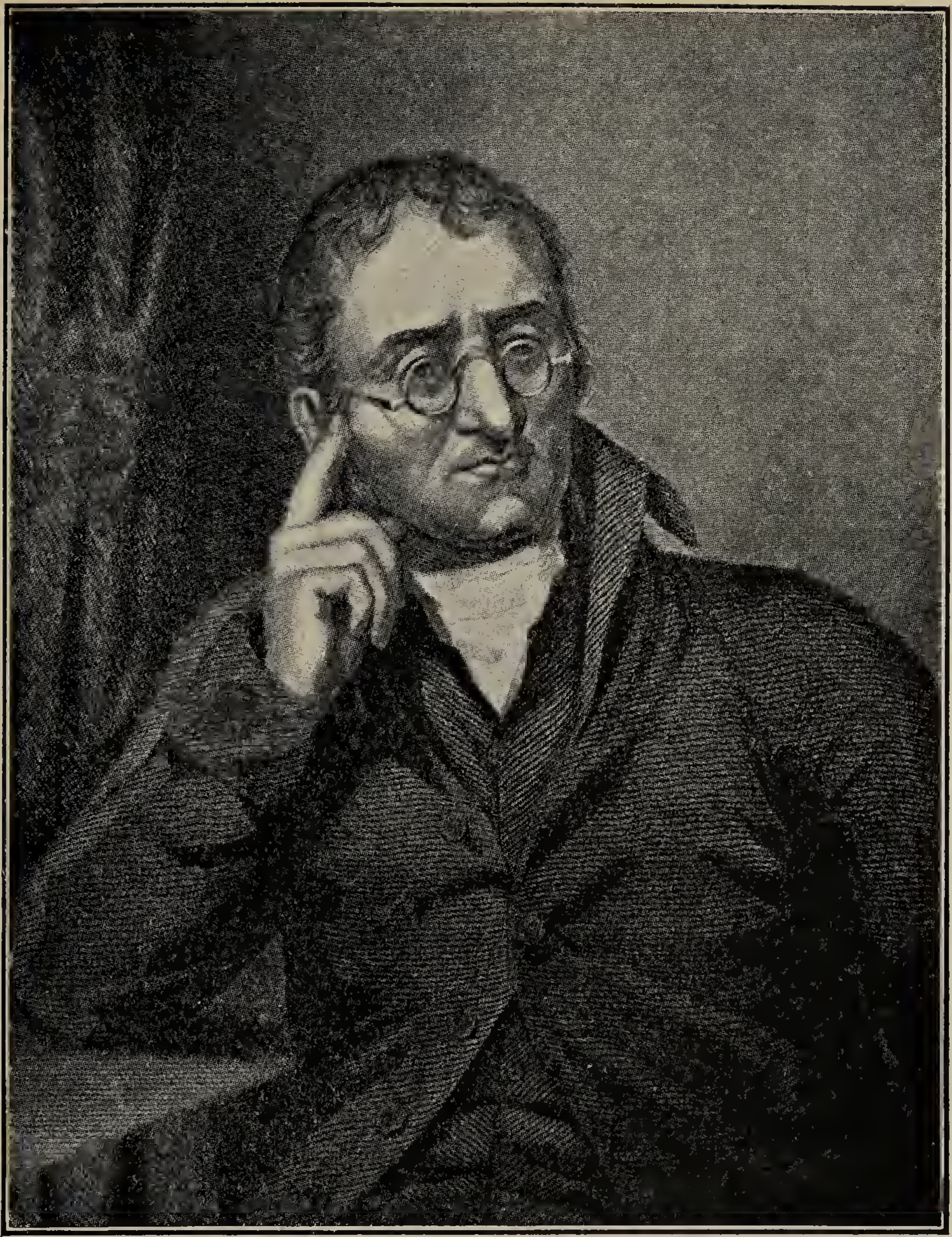
The Perry Pictures

Robert Boyle, 1626–1691

His book, "The Sceptical Chymist," is an early indication of modern scientific methods.

motion in liquids being less than in gases, and least of all in solids, although still present in some degree. The name molecules was given to the particles, and the theory was called the molecular theory.

By a quantitative analysis of compounds the proportions by weight of the elements composing them can be deter-



The Perry Pictures

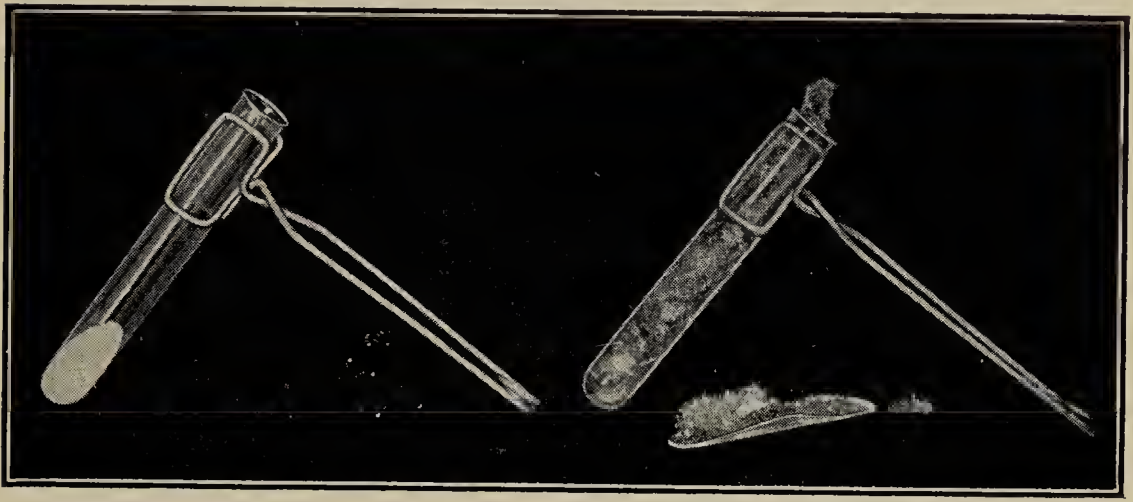
John Dalton, 1766–1844

The originator of the modern atomic theory.

mined. Such work established two most important truths called the law of definite proportions and the law of multiple proportions. The law of definite proportions states that every chemical compound contains always a definite proportion by weight of the elements composing it. The law of multiple proportions states that in special cases where two elements unite in different proportions to form different compounds, if a fixed weight of one of the elements be taken, the weights of the other element that combine with that fixed weight have a simple ratio to each other. In the case of hydrogen and oxygen, for instance, there are two compounds, water and hydrogen peroxides. Each of these contains always the same proportion of hydrogen to oxygen by weight and so each obeys the law of definite proportions. But the proportions are different in the two substances. In water the proportion of hydrogen to oxygen is 1 to 8, in hydrogen peroxide 1 to 16. Since 8 and 16 have a simple ratio, 1 to 2, these compounds obey the law of multiple proportions. Now as molecules have all the properties of their substances, these facts must be true of molecules, and to explain these facts our chemical theory must be extended and our conception of the composition of matter enlarged to include atoms.

The atomic theory states that molecules are made up of even smaller particles called atoms, these atoms not existing alone for any length of time but always combining quickly with other atoms to form molecules. In the case of each element the atoms are all alike, each molecule of the substance containing usually two or more exactly similar atoms. In the case of a compound, the atoms making up each molecule are those of several different elements; but whatever the composition of one molecule is, the others of the compound are of exactly the same composition. Atoms were for years considered indivisible, until the discovery of radio-activity by M. Becquerel and of radium by M. and Mme. Curie about the year 1898, and more recent work with radio-active substances proved the contrary for at least a number of elements.

It is probable that as new discoveries occur in this field the atomic theory will need further modification. At present for chemical purposes we still consider the atom the ultimate unit.

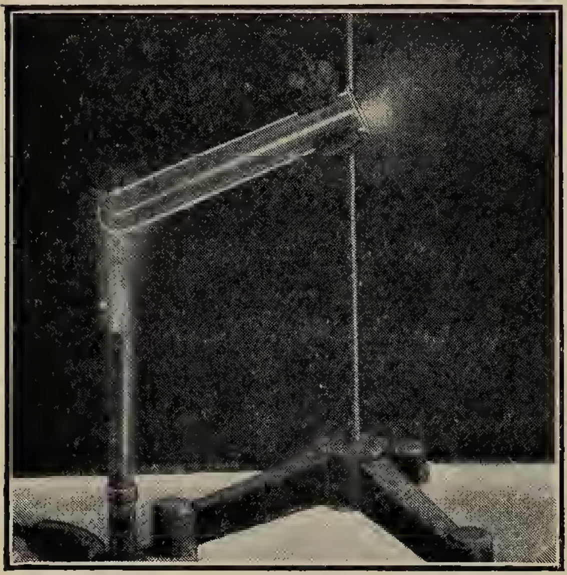


Sugar and the charred residue after heating it.

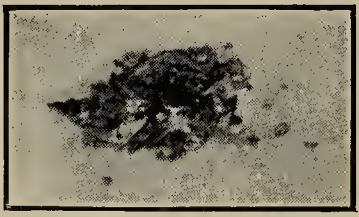
If chemical change be interpreted in terms of these two theories it may be said always to mean a breaking up of molecules. Refer to Experiment 1 in Chapter I, and see how simply this explains what happened there. In the case of sugar and magnesium the changes were chemical. The sugar molecules, large and complex, each composed of many atoms of oxygen, hydrogen and carbon, were broken up, and the atoms rearranged, so that finally by uniting with more oxygen from the air they formed new substances, water or hydrogen oxide, and carbon dioxide.

Magnesium in ribbon form, and magnesium oxide, or magnesia, the product formed when magnesium is burned.

A high temperature being needed for carbon to unite with oxygen, much carbon was left behind. With magnesium the process was simpler, the substance being an element, and the atoms of magnesium united quickly with those of oxygen in the air, giving out heat and light, and resulting in a white powder, magnesium oxide. With the iodine and ice no chemical changes occurred, because the molecules were not broken up but simply changed to gaseous or liquid form, that is, their rate of motion increased. The changes with the rubber rod and in the salt solution are more difficult to explain. Electricity enters into both of these cases, and we shall have to extend our theory still further to fit these circumstances. Before doing this we shall consider a convenient method for representing elements by letters, and extend it to the representation of compounds and reactions.



Iodine crystals sublimed, showing iodine vapor.



Iodine crystals.

DEFINITIONS

A **molecule** is the smallest particle of an element or compound that can exist alone and that has all the properties of the substance.

An **atom** is the smallest particle of an element and is, in general, indivisible.

QUESTIONS

1. Why do you think it is often necessary to heat substances before they will react?

2. Name instances in which it is evident that solids and liquids expand when heated and contract when cooled, and explain theoretically why this is so.

3. What important use is made of the force of gases heated under pressure?

4. Which contains most and which least energy of motion, ice, water, or steam?

5. Define synthesis, decomposition, replacement, double replacement, in terms of atoms and molecules, taking specific substances to illustrate.

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CHAPTER X

SYMBOLS

There is a short way of representing elements, compounds and reactions which not only saves time but makes it easy to calculate what quantities of substances will react with each other. This is by means of symbols, letters representing the elements. Usually the first letter, or first two letters, of the name are used, as H for hydrogen, O for oxygen, Br for bromine. Sometimes the first letters of the Latin name are used, as Fe for iron (ferrum), Ag for silver (argentum), and so forth. This symbol is by common consent taken to mean not only the substance, but one atom of the substance, and this makes it possible to represent with these letters molecules of elements and compounds. By ingenious applications of theory, borne out by experiments, it has been possible to determine how many atoms are present in the molecules of various substances, and so we have formulas, or groups of letters representing these molecules. Hydrogen, oxygen, nitrogen and chlorine are known to contain two atoms to the molecule. Their formulas are therefore written H_2, O_2, N_2, Cl_2, respectively. 2H would on the other hand mean two atoms of hydrogen not yet combined to form a molecule. Formulas for molecules of compounds are usually more complicated than these. A table of elements and their symbols will be found at the end of the book, and should be referred to at this point. Among the simpler formulas are HCl for hydrogen chloride, containing one atom of each element; CO_2 for carbon dioxide, containing one atom of carbon and two of oxygen; NaOH for sodium hydroxide, containing one atom each of sodium, oxygen and hydrogen; and so on.

The use of these symbols, with the plus sign for " and " or " with," and the arrow for " give," makes the writing of equations for reactions a much more simple and definite matter. Thus in place of saying

Carbon and oxygen give carbon dioxide,

we may put it thus:

$$C + O_2 \rightarrow CO_2.$$

Usually, however, there is need of balancing the equation by putting numbers before certain molecules. It would not do to write

$$H + O \rightarrow H_2O,$$

first because hydrogen and oxygen do not occur in single atoms but in molecules containing two atoms each. So we amend it to

$$H_2 + O_2 \rightarrow H_2O.$$

Here another difficulty is met. It has been found true that equal volumes of gases contain equal numbers of molecules. H_2 and O_2 stand, therefore, for equal volumes of these gases. Now it is further found by experiment that one volume of hydrogen does not combine with an equal volume of oxygen to make one volume of steam, but that to make steam with no hydrogen or oxygen remaining over, a volume of hydrogen combines with half that volume of oxygen to make a volume of steam equal to the volume of the original hydrogen. In the electrolysis of water, it will be remembered, half as much oxygen formed by volume as hydrogen. On the basis of oxygen this may be stated

2 volumes of hydrogen + 1 volume of oxygen → 2 volumes of steam.

This means a final revision of the equation. Instead of

$$H_2 + O_2 \rightarrow H_2O$$

write

$$2\,H_2 + O_2 \rightarrow 2\,H_2O,$$

which is the final and only correct form. This last change is seen to result in the same equation as that obtained by balancing the former one algebraically, and so in actual practice the modification may be made on that basis.

In writing equations for reactions, therefore, care must be taken that

(1) the substances actually react to give the products named.

(2) each substance be represented by its correct molecular formula.

(3) the equation be algebraically balanced by placing the necessary numbers before the molecules. These numbers affect the entire molecule. For instance, in $2H_2O$ we are dealing with four hydrogen atoms and two oxygen atoms.

There are evidently several factors in a reaction that are not expressed in such an equation. The conditions that cause it are not given, nor the nature of the substances, nor their energy content, nor whether the reaction is exothermic or endothermic. Arrows pointing up or down placed after a substance may show it to be a gas or a precipitate, as the case may be. Sometimes the number of calories of heat formed or used is given thus:

$$H + Cl \rightarrow HCl + 22{,}000 \text{ calories}$$

or

$$H + I \rightarrow HI - 6100 \text{ calories.}$$

An equation is an imperfect but condensed and very convenient expression of a chemical reaction.

The fact that the formula for hydrochloric acid is HCl and that for water H_2O, that the formula for marsh gas is CH_4 and that for carbon dioxide CO_2, leads to a consideration of what is termed valence. Why does one hydrogen atom combine with one chlorine atom, and two hydrogen atoms with one oxygen atom? Why does it need four hydrogen atoms to hold one atom of carbon, and only two atoms of oxygen for that same carbon atom? We express this by saying that the valence of chlorine is one, and of oxygen, two. The valence of a substance means the holding power of its atom, that of the hydrogen atom being considered one. The atoms of certain elements, such as chlorine and sodium, combine with or replace one atom of hydrogen, as the formulas HCl and $NaCl$ denote. The atoms of other elements, like oxygen and copper, combine with or replace two hydrogen atoms and make necessary such formulas as H_2O, CuO. Other elements have valences of three or four or even higher numbers. In cases where the law of multiple proportions applies, an element is said to have two or more valences; thus copper has a valence of two in cupric oxide, CuO, and of one in cuprous oxide, Cu_2O.

It is of course impossible to weigh atoms, since we cannot apprehend them in any way except by reason and imagination. It has, however, been found possible to discover their relative weights by taking some element as a standard and expressing the weights of other elements in terms of this one. Much confusion has been caused in the past by rival standards, some chemists preferring hydrogen = 1 as a standard and some preferring for greater convenience oxygen = 16. Oxygen is at present used as a standard by most chemists, and on this basis hydrogen has an atomic weight slightly greater than one. Hydrogen is the lightest substance known. Oxygen is about sixteen times as heavy, nitrogen fourteen times as heavy. Since, as has been said before, equal volumes of gases are believed to contain equal numbers of molecules, the fact that oxygen is sixteen times as heavy as hydrogen must mean that its molecule is sixteen times as heavy as the hydrogen molecule. Further, since each molecule of hydrogen and of oxygen contains two atoms, it is plain that the atom of oxygen is sixteen times as heavy as the atom of hydrogen, and that the molecule of oxygen is thirty-two times as heavy as the atom of hydrogen. These numbers, 16 and 32, are therefore called respectively the atomic and molecular weights of oxygen. Nitrogen, in like manner, has an atomic weight of 14 and a molecular weight of 28. The determination of atomic weights of other non-gaseous elements is a more complicated process, but it has been accomplished with fair accuracy.

If the number representing the weight of a molecule of a substance in terms of the hydrogen atom, as 32 for oxygen, 28 for nitrogen, 98 for sulphuric acid, be treated as a number of grams, we have what is called the gram-molecular weight of the substance. It is evident, then, that the following definition of the atomic weight of an element is possible:

The atomic weight of an element is the least weight of that element present in the gram-molecular weight of any of its compounds.

DEFINITIONS

A **symbol** is a sign standing for some word or idea. Symbols in chemistry are letters representing the elements.

A **formula** is a group of symbols representing a compound.

A **radical** is a group of atoms never occurring alone but forming part of a compound. These atoms remain together in many chemical reactions, and act as one atom.

Avogadro's hypothesis states that equal volumes of gases under like temperature and pressure contain equal numbers of molecules.

The **valence** of an element is the number of atoms of hydrogen that its atom combines with or replaces.

The **atomic weight** of an element is the weight of its atom in terms of an atom of hydrogen. The atomic weight may also be defined as the least weight of that element present in the gram-molecular weight of any of its compounds.

The **molecular weight** of a substance is the weight of its molecule in terms of the atom of hydrogen.

The **gram-molecular weight** is the molecular weight in grams.

QUESTIONS

1. With the aid of the table of elements and radicals, find the molecular weights of the following elements and compounds:

Oxygen, nitrogen, hydrogen, hydrochloric acid, sodium hydroxide, nitric acid, calcium hydroxide, carbon dioxide, sulphur dioxide, sodium chloride, water, hydrogen peroxide, magnesia, copper oxide, mercuric oxide, potassium chlorate, phosphorus pentoxide, ferric oxide, ferrous oxide, zinc chloride, silver nitrate, copper sulphate, lead nitrate, ferrous sulphate, potassium chloride, potassium bromide, potassium iodide, sodium sulphate, aluminum sulphate, aluminum hydroxide, ferric hydroxide, copper hydroxide, copper chloride, barium chloride, silver chloride, ferric nitrate, potassium hydroxide.

2. Place the following elements in columns according to their valences:

Hydrogen, oxygen, nitrogen, chlorine, sodium, calcium, carbon, sulphur, magnesium, copper, mercury, potassium, phosphorus, iron, zinc, silver, lead, bromine, iodine, aluminum, barium.

3. Take the reactions of Chapters V and VI, and write an equation in symbols, properly balanced, for each one, following the directions given in the text.

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CHAPTER XI

SOLUTIONS AND IONS

At the conclusion of the discussion of the atomic theory it was stated that the experiments in electrifying the rubber rod and in dissolving salt in water demanded an extension of this theory. This is a simpler matter now that the reactions involved can be expressed in the symbols explained in the preceding chapter. A clearer idea of what solution means is first necessary.

When salt is dissolved in water, what happens? How is this different from putting sand in water? It is evident that sand is not dissolved at all, but merely suspended, and may be recovered by filtering off the water through filter paper. Salt, on the contrary, disappears and distributes itself uniformly, in an invisible state, through the water. Indeed, substances in solution act very much as do gases, and obey the same laws. The exact nature of the change in the molecule is unknown; it evidently becomes transparent or translucent, that is, permits some light to pass through it; it retains its taste, odor and color, if it had any. We have already in a previous chapter given the name crystalline or crystalloid to such substances, because when they separate slowly from solution they do so in regular shapes called crystals.[1]

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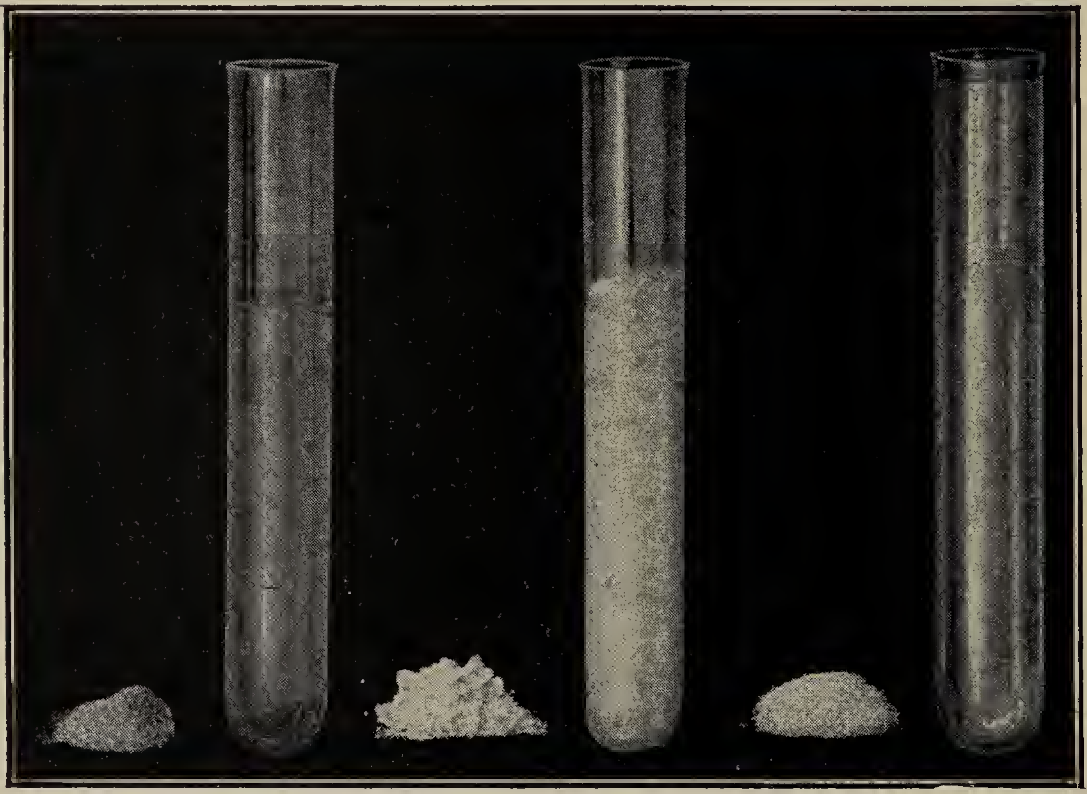
[1] For a short statement of the action of water as a solvent see A. Smith. Elementary Chemistry, p. 92.

Another large group of substances is called colloids, and includes starches, gums and other common substances. They are intermediate between insoluble substances like sand and soluble substances like salt. They neither dissolve completely in water nor may they be filtered out through ordinary filter paper. They are the subject of much important investigation at present. Many substances in our foods are colloids, and their digestion will be discussed in Part II.



1. Sand, and sand heated in water, showing an insoluble condition. 2. Starch, and starch heated in water, showing a colloidal condition. 3. Sugar, and sugar heated in water, showing a soluble condition.

Salt, like all the soluble acids, bases and salts so far worked with, belongs to a group that undergo a further change in water than mere solution, and an extension of the atomic theory has been necessary to explain their reactions. Recall the electrolysis of water and it will be remembered that water alone cannot be decomposed by the electric current. It is necessary to put something into it, like sulphuric acid, H_2SO_4, making a solution which will conduct the current. The conductivity of some substances in solution and other peculiarities in their action make it necessary to suppose that many of the molecules are divided or dissociated into two parts when the substance is dissolved, and that part of the chemical energy holding the atoms together in the molecule shows itself as charges of electricity on these two parts, one receiving

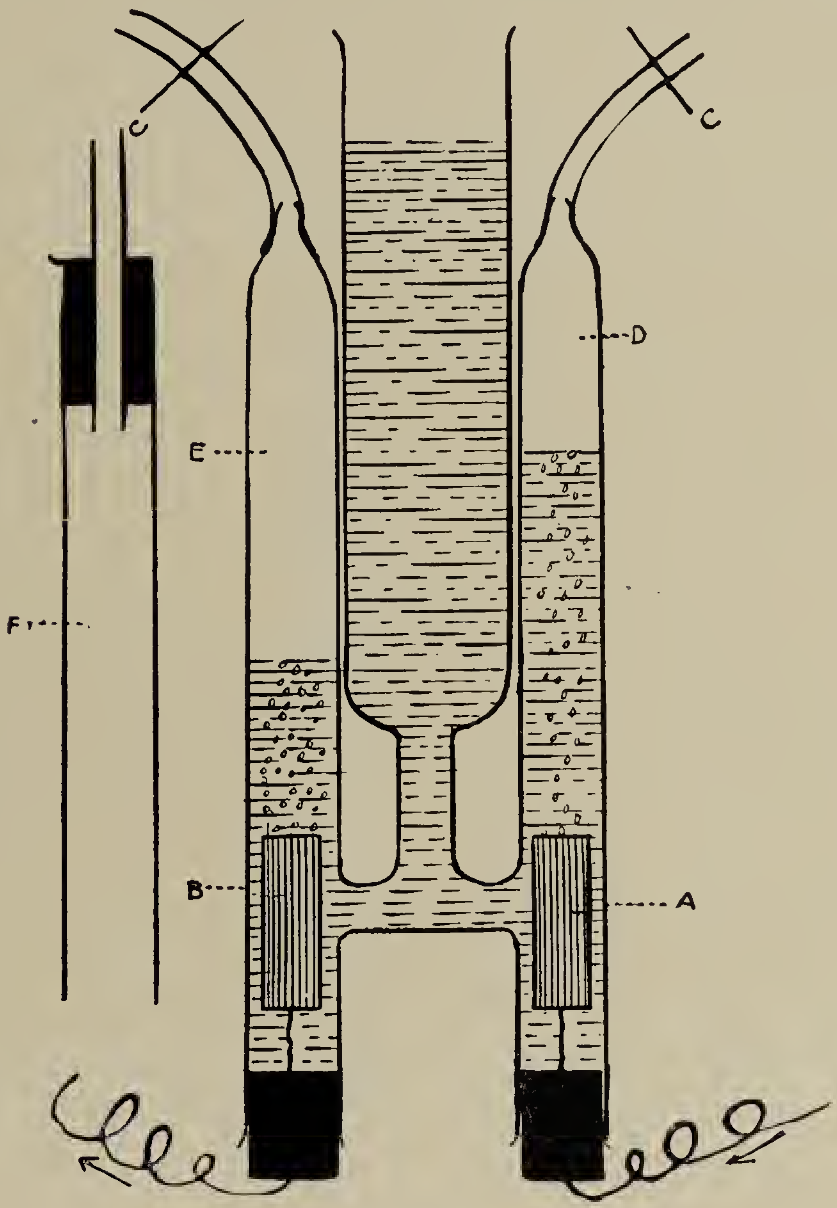


Diagram for Electrolysis of Water

a positive charge and the other receiving a negative charge.

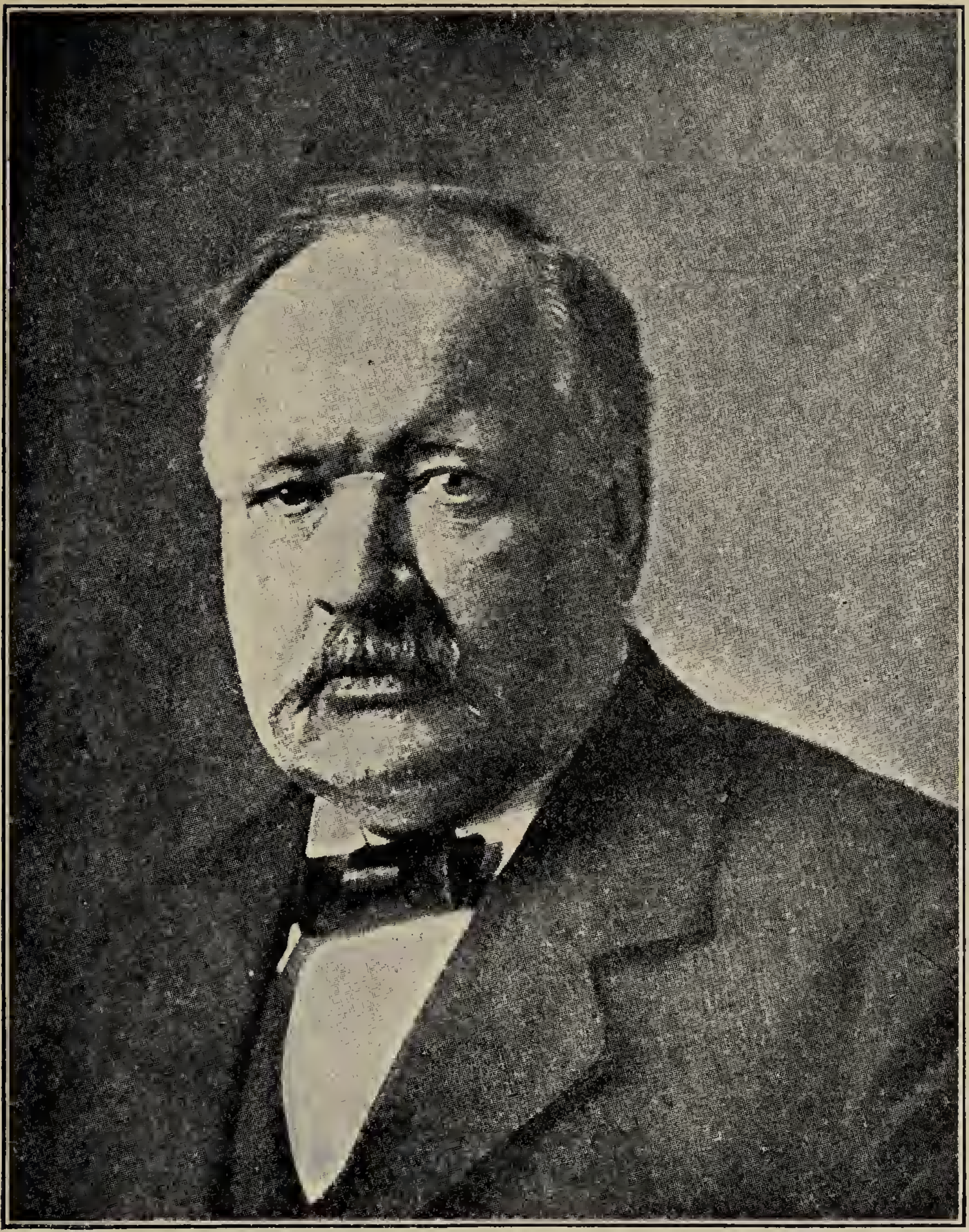
$$H_2SO_4 \rightarrow 2H^+ + SO_4^{--}$$

Each subdivision of the molecule is an atom or a group of atoms acting as one; each has, in addition, the electric charge which modifies its action. Hydrogen and the metals receive positive charges, while hydroxyl and acid radicals receive negative charges. When the electrodes are now introduced and the current turned on, the positive electrode where the current enters attracts the negatively charged SO_4^{--} groups until they touch it, much as the rubber rod attracted the pith ball. Then the two kinds of electricity neutralize each other, the particles lose their electric charges and are able to react with the water molecules to form hydrogen sulphate or sulphuric acid, setting free the oxygen which, it will be recalled, rose above the positive electrode.

$$2SO_4^{--} + 2H_2O \rightarrow 2H_2SO_4 + O_2\uparrow$$

At the negative electrode the positive hydrogen particles collect, their charges are neutralized, and they become atoms and then join to form molecules, which rise above the electrode. Sulphuric acid is constantly being re-formed at the positive electrode, and so the reaction might be imagined to go on until all the water had disappeared and only sulphuric acid remained.

These particles carrying electric charges are called ions, the substance in solution is said to be ionized, and the theory is called the ionization theory, or the theory of electrolytic dissociation. Ionization varies with circumstances. Some substances, such as sugar and glycerine, are not ionized at all; others are more or less completely ionized, depending on the degree of concentration, nature



From "Elementary Household Chemistry"
By John Ferguson Snell, published by The Macmillan Co.

Svante August Arrhenius

The Swedish scientist who in 1887 originated the modern theory of ionization of electrolytes.

of the solvent or substance in which they are dissolved, and so on. The more concentrated a solution is, the less the solute, or substance in solution, is dissociated. Therefore in reactions which are ionic, or which take place between substances in solution, it is better to use dilute solutions than concentrated. Because substances that are highly ionized in solution are good conductors of the electric current, they are called electrolytes, and other substances non-electrolytes. It must be emphasized, however, that the passage of the electric current does not cause the dissociation and the charging of the ions, but that the reverse is true, and that the dissociation caused by the solvent permits the passage of the current because of the charges of electricity already carried by the ions.

Water as a solvent causes a high degree of dissociation. This is significant in view of the fact that water is present in almost all chemical reactions. It is possible that all chemical reactions are of an ionic nature and so, in the last analysis, electric. At present we confine the term ionic to those reactions taking place in solution.

Color changes in solution are always fascinating, and it is interesting to know that these may often be interpreted in terms of the colors of molecules and ions. Take, for instance, the case of copper bromide. The substance is of a dark brown color, and this is also the color of its concentrated solution, bearing out the statement that in concentrated solution a substance is still in its molecular state. As water is added very slowly the color changes, first to a bottle green, and then through lighter shades of green until finally, when the solution is very dilute, a pale blue color results. Since it is known from solutions of other bromides that the bromine ion Br^-, is colorless, the changes of color must be due to varying proportions of the dark brown molecule and the copper ion, Cu^{++}. The latter is pale blue, and therefore as dilution proceeds the increasing number of dissociated molecules adds the light blue of the copper ion in greater proportion to the color scheme, and so with the dark molecules there is produced first a dark green, then lighter greens, until finally almost complete dissociation prevails, and the color of the dilute solution becomes that of the copper ion alone.

$$CuBr_2 \rightarrow Cu^{++} + 2Br^{-}.$$

This last equation illustrates the fact that the number of electric charges carried by an ion is considered to be equal to its valence, and that one copper ion carrying two positive charges, Cu^{++}, will balance or neutralize two bromine ions, each carrying one negative charge, $Br^{-} + Br^{-}$, or $2\,Br^{-}$.

It will be interesting here to recall the subject of acids and bases and their effect on each other, called neutralization. In Chapter VI acids were stated to contain hydrogen which could be replaced by certain metals, and bases to contain the radical hydroxyl, OH. It was stated that some acids and bases were more active than others. This may now be explained as due to the fact that in solution more of their molecules are dissociated. The more active or stronger an acid is in solution, the more hydrogen ions it contains. When solutions of acids and bases are put together, their hydrogen and hydroxyl ions unite to form molecules of water.

$$H^{+} + OH^{-} \rightarrow H_2O.$$

This is because water is not dissociated to any extent, and when hydrogen ions and hydroxyl ions occur in the same solution, they always unite to form molecules. The characteristics of both acids and bases disappear when their solutions are mixed, because these characteristics are dependent on hydrogen and hydroxyl ions. The effect of acids on litmus, their sour taste, and their reactions with zinc and iron to form salts and liberate hydrogen, are all due to the presence of the hydrogen ion, H^{+}. Likewise the action of bases in changing red litmus blue and their slippery or soapy feeling depend on the presence of the hydroxyl ion, OH^-. These properties all disappear, therefore, when the ions unite to form water. A typical set of ionic equations may be written thus:

$$HCl \rightarrow H^+ + Cl^-.$$
$$NaOH \rightarrow Na^+ + OH^-.$$
$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O.$$

Such salts as sodium chloride, formed by the neutralization of a strong or highly dissociated base, and a strong acid, are normal salts, and their solutions are neutral to litmus. . Other salts have been described in Chapter VII as alkaline or acid in solution. How can this be explained by the dissociation theory?

These salts are all found to be formed by the neutralization of a strong acid by a weak base, or a weak acid by a strong base. Copper sulphate, ferric chloride, lead nitrate are illustrations of acid salts, and it will be observed that sulphuric, hydrochloric and nitric acid are strong acids, while copper, iron and lead hydroxides are weak. On the other hand sodium carbonate, potassium sulphide, ammonium acetate, are alkaline salts, formed by the union of the strong bases sodium, potassium and ammonium hydroxide with the weak acids carbonic and acetic, and hydrogen sulphide. Such phenomena are explained as due to the action of the few ions of water that exist, and may be termed the reverse of neutralization. Instead of having a base and an acid producing a salt and water, we have a salt dissolved in water producing a base and an acid, one of the two undissociated. Symbols will make this plainer.

When sodium chloride is dissolved in water we have the reaction

$$NaCl \rightarrow Na^+ + Cl^-.$$

It must be imagined that there exist a few molecules of water dissociated into ions, H^+ and OH^-, so that the four kinds of ions exist together in the solution. Because both sodium hydroxide and hydrochloric acid are highly dissociated substances, these ions do not unite to form molecules of either acid or base, and the solution is neutral to litmus because it does not contain a preponderance of either hydrogen or hydroxyl ions. When sodium carbonate is dissolved in water, it also dissociates into ions

$$Na_2CO_3 \rightarrow 2Na^+ + CO_3^{--},$$

and there are also the few hydrogen and hydroxyl ions of water present. But because carbonic acid is a weak acid, or not highly dissociated, the hydrogen ions and the carbonic radical ions do unite to form molecules of the acid

$$2H^+ + CO_3^{--} \rightarrow H_2CO_3,$$

leaving in the solution an excess of hydroxyl ions with the sodium ions. It is the presence of these hydroxyl ions which gives the solution its alkaline reaction. Likewise when copper sulphate is dissolved and dissociated, the copper ions unite with the few hydroxyl ions of water present to form undissociated copper hydroxide, leaving hydrogen ions in solution, and so giving an acid reaction. This action between a substance in solution and the ions of water is called hydrolysis, and forms a very important feature of the chemistry of foods and cleaning. It will be referred to again in Part II.

The ionization theory makes it easier to picture what happens during replacement reactions. When chlorine replaces iodine in its solution of potassium iodide, for instance, the chlorine atom is pictured as taking over the negative charge from the iodine ion, thereby becoming ionic, and the iodine becomes molecular and may be shaken out, uncombined, by dissolving it in carbon disulphide. Ions play a part in the case of the displacement by a metal of another from its salt, as in Experiment 12, between iron and copper sulphate, or of hydrogen by a metal, as with hydrochloric acid and zinc. In the case of the metals and hydrogen, experiments make it possible to arrange these elements in a table which may be called the *displacement series*, found in Appendix B according to which a metal may be displaced from a solution of its salts by any metal coming before it in the series, and may in its turn displace any one following it. In these instances, as with chlorine and bromine, we think of the reaction as consisting in a transfer of electric charge from one ion to the atom of another element, which leaves the first element in an atomic or molecular condition. Thus for hydrochloric acid and zinc the reaction may be written

$$2H^+ + 2Cl^- + 2Zn \rightarrow Zn^{++} + 2Cl^- + H_2 \uparrow$$

and for iron and copper sulphate

$$Cu^{++} + SO_4{}^{--} + Fe \rightarrow Fe^{++} + SO_4{}^{--} + Cu.$$

DEFINITIONS

A **solution** is a uniform mixture, usually of a solid and a liquid. The solid is called the solute, and the liquid the solvent.

A **colloid** is a substance intermediate in nature between soluble and insoluble substances.

An **ion** is an atom or a radical carrying an electric charge.

An **electrolyte** is a substance whose solution is a conductor of the electric current. Its molecules are dissociated into ions by the solvent.

An **acid** is a substance whose solution contains hydrogen ions.

A **base** is a substance whose solution contains hydroxyl ions.

Neutralization consists in the union of hydrogen and hydroxyl ions to form molecules of water.

Hydrolysis is interaction between a substance and the hydrogen and hydroxyl ions of water. It is the reverse of neutralization.

QUESTIONS

1. Why does water not conduct the electric current?

2. Could you electrolyze water containing sugar in solution? Give reason for your answer.

3. Why was dilute, not concentrated, hydrochloric acid used on zinc to produce hydrogen?

4. Why was the blue litmus paper moistened before being held at the end of the delivery tube when hydrogen chloride was being formed in Experiment 6?

5. Why does hydrochloric acid on copper not produce hydrogen when it does on zinc? Judging from this fact, would copper displace zinc from its salts, or vice versa?

6. Take the reactions written in Question 3 of Chapter X, and rewrite them as ionic equations.

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CHAPTER XII

REVERSIBLE AND COMPLETED REACTIONS

It will be an easier matter now to understand when reactions will take place. When two substances such as potassium bromide and chlorine are put together, the greater heat of formation of potassium chloride determines its formation, and bromine is liberated, while bromine and potassium chloride would not react. When copper is placed in silver nitrate, silver is liberated because of the greater chemical activity of the copper, while silver and copper nitrate would fail to react. Reactions are thus seen to be largely dependent on the energy content of the compounds and elements involved. In general, no reaction will take place between substances unless some special cause exists, and it is well to consider many reactions as reversible unless there is good reason for them to complete themselves.[1] For example, let us take the reaction between iron and steam. If these substances are sealed into a tube and heated, a certain amount of iron oxide and hydrogen are formed, but after a time it is noticed that no more of these substances is formed, and yet that some iron and steam are still present. What has happened is that some iron oxide and hydrogen when formed immediately interact again to form iron and steam, until finally a state of dynamic equilibrium is established, in which nothing seems to be happening, but where, in reality, two processes are taking place, perfectly balanced against each other. This may be stated thus:

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[1] Of course many reactions cannot be reversed under any conditions; as, for instance, the decomposition of potassium chlorate. The decomposition of mercuric oxide, on the other hand, is reversible.

$$3Fe + 4H_2O \rightleftarrows Fe_3O_4 + 4H_2$$

reading the double arrow "are in equilibrium with."

Reactions are reversible not only between different substances but under other conditions, such as between different physical states of the same substance, water and steam, for instance; between the portion of a substance in solution and the portion still undissolved; and between the undissociated molecules of the substance in solution and the ions of the dissociated molecules. Interaction is constantly going on in all these cases, faster in one direction than in the other, until the reacting masses reach a condition, not necessarily of equality, but such as the circumstances of temperature, and so forth, demand at the time. When this state is reached we say equilibrium is established. Interaction still goes on, however, as many changes taking place in one direction as in the opposite direction. As many molecules of water become steam as of steam become water; as many molecules of the still undissolved portion go into solution as of the dissolved portion separate out of solution; as many molecules in solution dissociate into ions as pairs of ions unite to form molecules. This is called dynamic equilibrium, because while there is balance there is also constant change. We shall understand all evident reactions better if we realize this ceaseless interaction under apparently inactive conditions.

Someone has compared such conditions of constant activity to the great wave seen at Niagara Falls. Close at hand it is evident that tremendous movement is taking place in the rushing water. At a great distance the crest of the wave can still be distinguished, but now it seems stationary, all its energy is lost and only its permanent form remains. One might say, if one had never seen it closer, that there could be no motion in such a stable thing. So we, miles distant, as it were, from the ultimate particles which make up matter, may find it hard to believe in its unceasing movement. Perhaps if we remember the wave of Niagara, this will seem easier.

A chief reason why many reactions are not reversible is because in some way one of the products formed escapes from the field of action. In the reaction between iron and steam already discussed, if the hydrogen escapes unchecked and steam is constantly fed into the tube the action is forced to proceed in one direction only until no more iron remains.

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2\uparrow.$$

This principle is called the law of mass action, and may be stated thus:

The speed of a chemical reaction may be increased in one direction by increasing the active mass of one of the substances that produce the action in that direction.

We constantly increase the active mass of steam and diminish the mass of hydrogen by allowing it to escape. Therefore in endeavoring to maintain equilibrium the reaction completes itself in one direction. Had we started with iron oxide and hydrogen, and allowed the steam formed to escape, we should have had the opposite reaction.

$$Fe_3O_4 + 4H_2 \rightarrow 3Fe + 4H_2O\uparrow.$$

There are three chief ways in which a substance may leave the field of action and so cause a reaction to go to an end — by volatility, by insolubility, or by non-ionization. One instance of volatility has already been cited. Another is the action of sulphuric acid on sodium chloride in the making of hydrogen chloride, when, at the temperature used in the experiment, hydrogen chloride is volatile while the other compounds are not. Numerous cases of insolubility determine reactions among solutions. In parts (*a*) and (*b*) of Experiment 12 sodium hydroxide and ferric nitrate give an insoluble precipitate of ferric hydroxide. In the solution are sodium and hydroxyl ions, iron and nitrate ions, and molecules of sodium hydroxide, ferric nitrate, sodium nitrate and ferric hydroxide. But, of these, ferric hydroxide is insoluble, which means that instead of remaining in solution its molecules separate out and leave the field of action, which is, in this case, the solution. This disturbs the equilibrium of the particles in solution, and to readjust it more ferric and hydroxyl ions unite to form molecules. These continue to separate out and more continue to be formed until the reaction is completed in that direction. With sodium hydroxide and potassium nitrate we had the reversible reaction

$$NaOH + KNO_3 \rightleftarrows NaNO_3 + KOH,$$

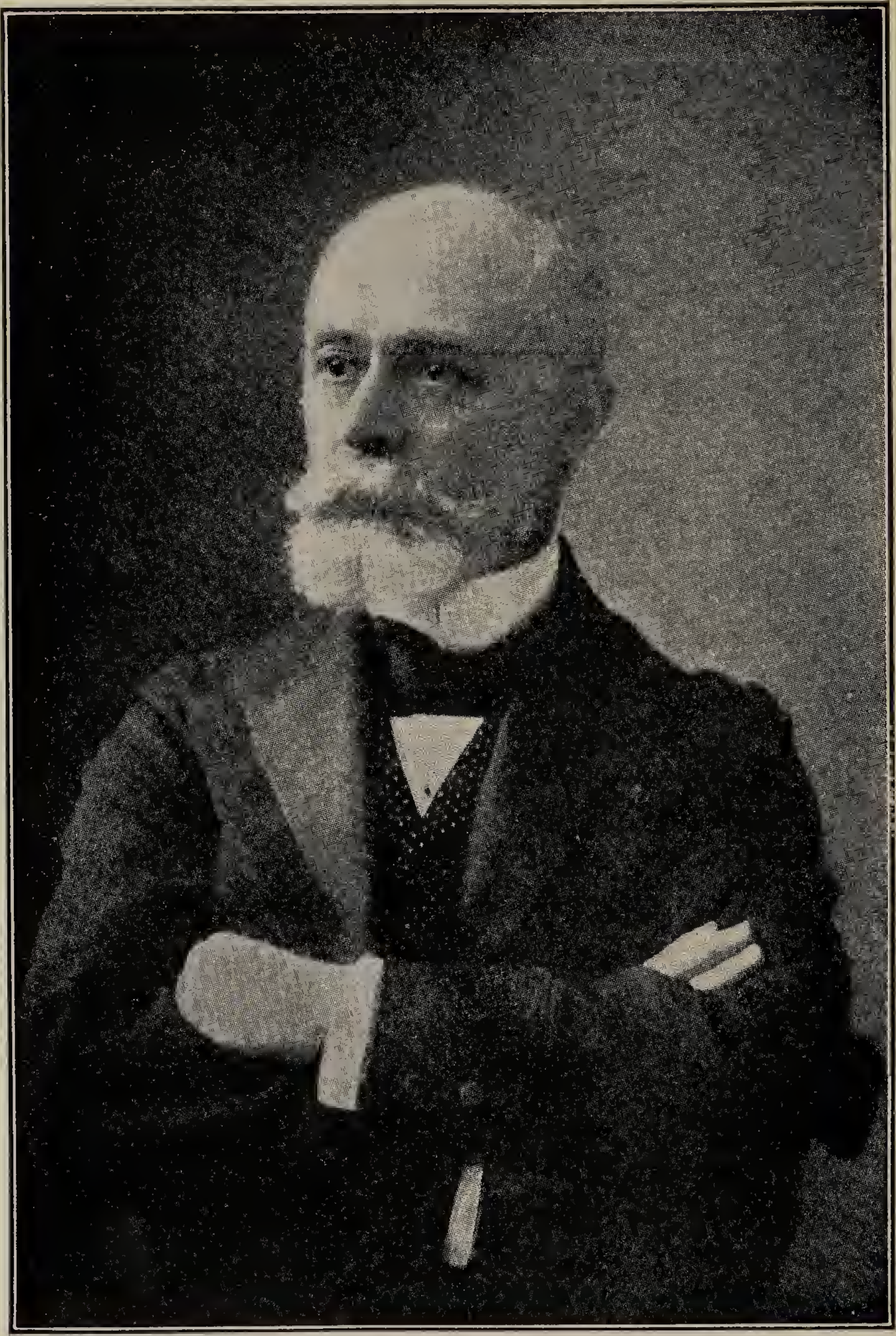
and nothing appeared to happen. But in the other case we had

$$3Na^+ + 3OH^- + Fe^{+++} + 3NO_3^- \rightarrow 3Na^+ + 3NO_3^- \\ + Fe(OH)_3 \downarrow,$$

and the reaction was completed.

Since reactions in solution are ionic, if a substance is not ionized or dissociated into ions no reaction between it and other substances takes place. Or if such a substance may be formed as a result of mixing solutions, this substance is removed, by its undissociated condition, from the field of action, and the reaction is completed as in the cases of volatile and insoluble compounds. Water is such a substance, very slightly dissociated into hydrogen and hydroxyl ions. Its formation explains neutralization of an acid and a base to form a salt. For instance,

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O,$$



From "Elementary Chemistry"
By Alexander Smith, published by The Century Co.

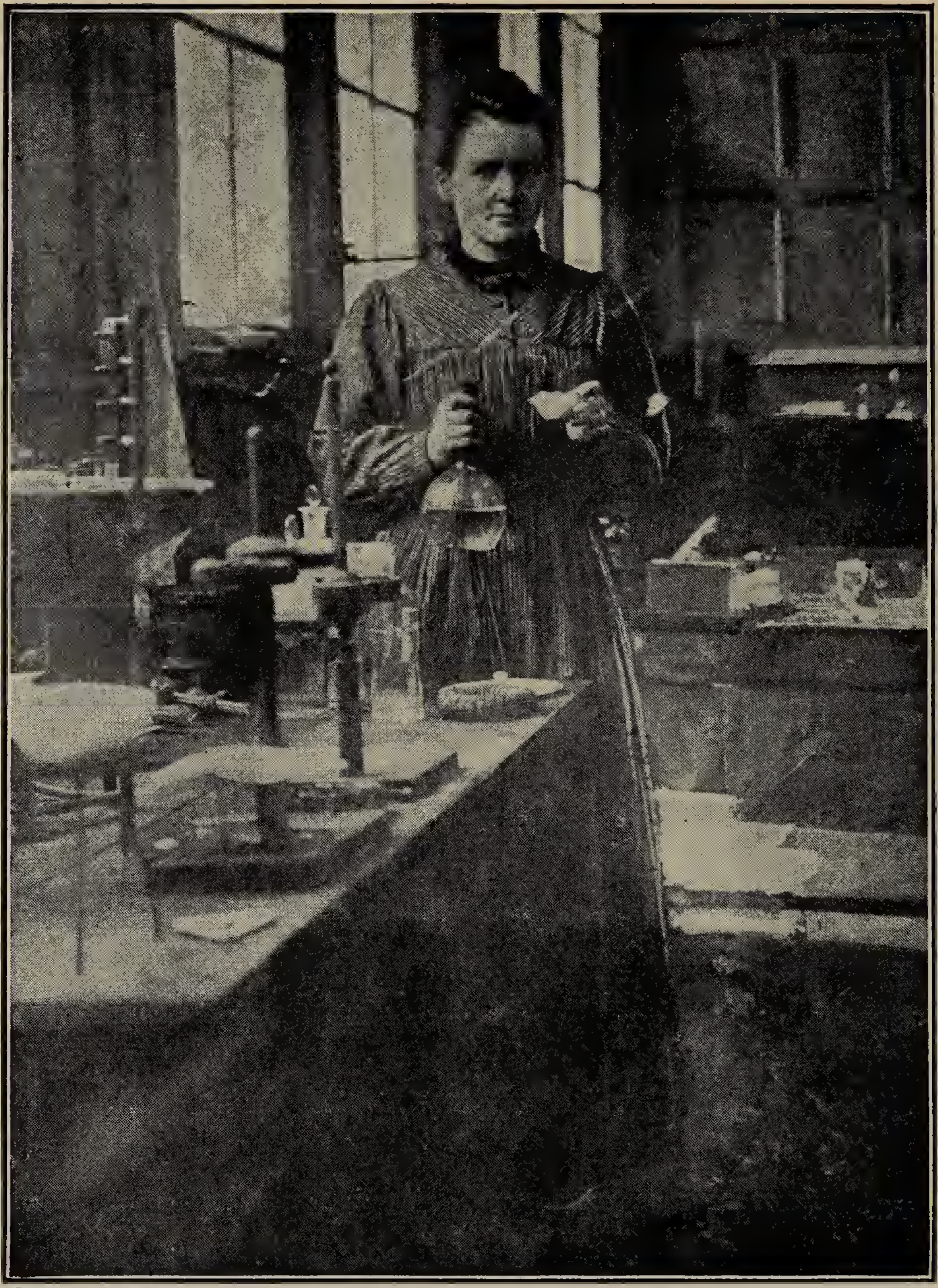
Henri Becquerel, 1852–1908

French physicist. Discoverer of radio-activity.

a reaction not reversible because of the almost complete lack of ions of water.

It has been seen that in order to understand under what conditions chemical changes take place some study of theory is necessary. This theory states that matter is composed of infinitely small atoms, linked together, by a force called chemical energy, into larger aggregates called molecules. Any chemical change means a breaking apart of molecules and a re-arrangement and re-uniting of the atoms. In solution, the molecules of many substances undergo a subdivision into ions, particles carrying electric charges, and reactions in solution take place between these ions. Reacting substances often remain in a state of dynamic equilibrium in which the speed of action in opposite directions is equal. Reactions go to an end without the application of external energy when it is possible to form a substance with a greater heat of formation than the other substances reacting; when a metal may be driven out of its salt by another metal of greater chemical activity; or when the law of mass action, under the conditions of the experiment, results in the removal, by reason of volatility, insolubility, or non-ionization, of some product from the field of action.

The atomic theory has received startling confirmation, if some modification, at the hands of recent investigators into the region of radioactivity. Certain elements of heavy atomic weight, such as uranium and radium, are now known to be constantly and naturally undergoing decomposition into simpler forms such as lead and helium. Chemists have not yet mastered the process so as to control it, but the undisputed facts of such change have led to a wonderful new way of looking upon all forms of matter, all kinds of atoms, as composed of many exactly similar corpuscles or electrons, and viewing the differences be-



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Mme. Curie in her Laboratory

The discoverer of radium.

tween elements as due to differences in number and arrangement of these little particles. This is called the corpuscular or electrotonic theory of the composition of matter, and in so far as it explains many hitherto inexplicable phenomena, and succeeds in giving us a more unified explanation of the universe, it is so much the more interesting and satisfying.

DEFINITIONS

A **reversible reaction** is one that may proceed in either or both of two opposite directions according to the conditions present.

Dynamic equilibrium is the condition of balance which exists when the two actions that make up a reversible reaction are proceeding at equal speed.

The **law of mass action** states that the speed of a chemical reaction may be increased in one direction by increasing the active mass of one of the substances that produce the action in that direction.

QUESTION

Review the experiments thus far performed and completed, and give a reason why each was completed.

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DALTON COLLECTING MARSH FIRE GAS

From a painting by Ford Madox Brown, in the Town Hall, Manchester
By permission of the Town Hall Committee of the Manchester Corporation

PART II

CHAPTER I

CARBON COMPOUNDS

In the first part of the book many different substances and classes of compounds have been mentioned. Some of them occur in nature as such, and many are prepared by man's ingenuity from substances found naturally. If the source of all these substances be considered we find that they all originate from the rock material of the earth, the air and water. This is true not only of such substances as sodium chloride and sulphuric acid, for instance, which are evidently mineral or inorganic, but of such substances as sugar and alcohol, which are found in plant and animal bodies, or in organisms, and are made in nature only as a result of the activity of some living thing. In the last analysis these also have their origin from soil, water and air. The element carbon found free in nature as coal, graphite and diamond, and combined in limestone and in a gas which makes up a small part of the air, characterizes all organic compounds. Because carbon forms so many compounds the chemistry of these compounds is separated from the chemistry of other compounds and is called organic chemistry.

The term "organic" was given at a time when these compounds were thought to be produced only by living bodies. Now many of these same substances can be made in the laboratory due to man's increased knowledge, and they are identically the same substances as made by the living plant or animal. A good illustration of this is the vanilla extract so familiar to us. To get this flavoring, beans from the vanilla plant may be used giving the characteristic odor and taste; also oil of cloves may be so treated chemically that the resulting substance is the same as that obtained from the vanilla bean. Chemical knowledge has so increased in late years that many substances which used to be made from plants or animals are now made in the laboratory synthetically, as it is called, much cheaper and in much larger quantities. This condition of affairs for dyestuffs and many other substances was existent before the Great War, and we are destined to see further and further development along these lines.

The best way to obtain first-hand knowledge of the chemical composition of these substances is to burn some of them in air and to observe the residues and the oxides formed.

EXPERIMENT 13

Purpose.

To study some of the substances formed when some common organic substances are burned.

Directions.

Burn a few c. c. of kerosene in a watch crystal, and hold over the flame a cool, dry bottle to catch any products of the burning. Note the appearance of the bottle.

Repeat with denatured alcohol.

Ignite a few splinters of wood in an iron dish, and hold bottle as before. Heat some splints in a test-tube in the flame. Describe reaction. From time to time place the mouth of the tube to the flame. Result?

When the tube is cool, scrape it out or break it open. Describe contents.

Repeat with sugar.

Hold a bottle over a burning candle.

Powder a little soda-lime and mix with bits of horn. Heat in a test-tube, placing at the mouth moist pink litmus and turmeric paper. Notice reaction in tube and appearance of paper, and odor. Examine contents of tube when cool.

Questions. What elements were present in all the compounds burned? How was their presence shown? What other element must have been present in the horn to give ammonia?

EXPERIMENT 14

Purpose.

To study one of the gaseous products formed when some common organic substances are burned.

Directions.

a. Fasten a piece of wire around a bit of charcoal. Hold it in Bunsen burner flame until it is burning, then let it burn in a wide-mouthed bottle one-third full of lime water. (Charcoal does not burn steadily, so repeat until a change occurs in the lime water.)

b. Lower a burning candle into a wide-mouthed bottle or jar. Let it burn a few minutes. Then add lime water and shake.

c. Burn some sugar in a deflagrating spoon in a bottle and proceed as in (*b*).

d. Hold a burning splinter in a bottle, then proceed as in (*b*).

e. Put enough alcohol into a wide-mouthed bottle to cover the bottom, then set fire to it. When completely burned, proceed as in (*b*).

f. Into a test-tube one-third full of lime water blow the gas from the lungs by means of a glass tube.

g. Into a test-tube put a few pieces of marble. Add hydrochloric acid. Quickly insert a rubber stopper fitted with an exit tube. Put this exit tube into a tube half-full of lime water.

Questions. What was the action in the lime water in all this experiment? If carbon is an element, and was burned in (*a*), what was the compound formed? Was the same compound formed in all cases?

Write in symbols the equation for (*g*). (This is the usual method for preparing this gas.)

Comparing with preceding experiment, what two oxides are formed when organic substances are burned?

The carbon compounds are much more numerous than those of any other element, and many of them are much more complex. That again makes a division into the inorganic and organic groups seem necessary. Most of these compounds are combinations of carbon, hydrogen and oxygen, as we have seen in the experiments, though some of them contain nitrogen and some of them contain sulphur, phosphorus, iron, potassium, sodium and a few other elements in small quantities.

Carbon compounds may be divided into three large groups: hydrocarbons, containing only carbon and hydrogen, and their derivatives, the alcohols, acids, esters and so forth; carbohydrates, containing oxygen in addition to carbon and hydrogen; and proteins, containing carbon, hydrogen, oxygen, and in addition nitrogen and sulphur with small quantities of a few other elements.

The simplest of the hydrocarbons is methane, or marsh gas, CH_4. It is found in marshes where vegetable matter is undergoing decay. It is also one of the important constituents of the illuminating gas made by the destructive distillation of soft coal. Methane is the lowest of a series of hydrocarbons, which is known as the paraffin series because paraffin is a mixture of several of the higher members of the series. They are found naturally in petroleum, naphtha and the natural waxes. The lower members of the series are gases, the next are liquids and the higher ones are solids.

Many hydrocarbons, if arranged in order according to increasing molecular weight, are found to be similar in constitution and in certain chemical properties, but the physical properties differ. Such is true if one begins with methane, CH_4. The next is ethane C_2H_6, the next is propane C_3H_8, the next is butane C_4H_{10}, and so on. Each one differs from the preceding by CH_2, so that the formula for any member is C_nH_{2n+2}. The compounds of this series are very stable substances and are said to be saturated. This means that the valency of carbon is four, and new substances made from any one of these compounds are always made by the replacing of hydrogen atoms, not by the addition of any element to the atoms already present. If for instance methane be treated with chlorine it is possible to replace any or all of the hydrogen atoms by chlorine atoms. When three atoms are replaced chloroform ($CHCl_3$), so commonly used as an anæsthetic, is formed. If all four hydrogen atoms are replaced carbon tetrachloride (CCl_4), which is the basis of the cleaning fluid carbona, is formed. If three atoms of hydrogen are replaced by three atoms of iodine, iodoform (CHI_3) the disinfectant is formed.

Ethylene (C_2H_4) is the beginning of another series of hydrocarbons whose general formula is C_nH_{2n}. They are unsaturated compounds and therefore have properties very different from the paraffin series. They are capable of forming additive products.

Acetylene (C_2H_2), the familiar illuminant, is the lowest member of another series bearing its name. The general formula is C_nH_{2n-2}.

If methane is oxidized properly one of the hydrogen atoms is replaced by an hydroxyl (OH) group. This process, with other compounds of the series, gives compounds with the general formula of $C_nH_{2n+1}OH$, and which bear the name of alcohols. Since there are many members of this series and each member may have several replaceable hydrogen atoms, the number of alcohols is very great. Wood and grain alcohols (CH_3OH and C_2H_5OH) represent two of the most familiar mono-hydric alcohols, while glycerine $C_3H_5(OH)_3$ is a tri-hydric alcohol. These alcohols resemble the metallic hydroxide of inorganic chemistry in composition, but are not dissociated and do not affect litmus. In the same way the ethers resemble the metallic oxides. The ethers contain two hydrocarbon radicals joined to one oxygen atom. The most common one is ethyl ether $C_2H_5OC_2H_5$, the liquid so commonly used as an anæsthetic.

Aldehydes like formaldehyde HCHO are derived from the monohydric alcohols and have as the characteristic radical CHO.

The fatty acids are derived from the aldehydes having the CHO group replaced by the COOH radical. The term " fatty " is applied because so many of them occur in natural fats, and resemble fats in physical properties. Acetic acid CH_3COOH, commonly known as vinegar, palmitic $C_{15}H_{31}COOH$ and stearic acids $C_{17}H_{35}COOH$, are familiar examples of this class of compounds. When alcohols combine with fatty acids the resulting substance is a salt, comparable to the inorganic salts. Such compounds are called ethereal salts. These ethereal salts are found often in flowers and fruits, so that the artificially prepared substance may be used as a flavoring or perfume.

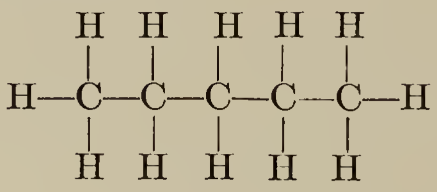
Ethyl acetate $CH_3COOC_2H_5$ is known as " banana oil," while anyl acetate $CH_3COOC_5H_{11}$ is known as " pear oil." If the salt is made from a metallic hydroxide and a fatty acid it is known as an ester. When the chapter on soap is studied this distinction will be clearer.

The organic substances which have been mentioned are only a few of the simpler fatty compounds. Starting with benzene C_6H_6 another group of organic compounds may be built up. These differ in several ways from the ones built up from methane or marsh gas, and are known as the aromatic compounds. To these belong the phenols, aromatic alcohols, ketones, aldehydes and dyes. The fact that series of compounds can be built up from these simple ones indicates the complexity of the subject. Again, when other elements like sulphur, chlorine and iodine enter into the composition a still better idea of the extent of organic chemistry is gained.

The carbohydrates and proteins, unlike the hydrocarbons, are compounds whose molecular composition is not so well defined as that of the alcohols and the rest, but they are compounds which occur naturally in large quantities and which form a large part of our food. The carbohydrates include the starches, sugars and cellulose. Starches and sugar are familiar enough to need no comment here. The celluloses are the substances which make up the cell walls of plant tissue and which, in the form of paper, cotton and wood, are of utmost importance, especially in war times.

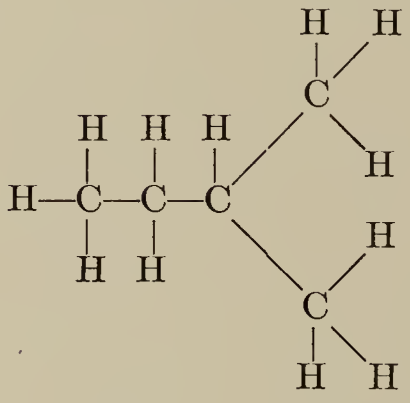
One of the most interesting facts about organic compounds is that there may be several compounds having the same percentage composition but having very different properties. Cane sugar or sucrose, malt sugar or maltose, and milk sugar or lactose all bear the formula $C_{12}H_{22}O_{11}$, but their physical properties differ considerably as well as their occurrence. Cane sugar is found in plants, milk sugar is found in animals only, while malt sugar is made by the action of malt on starch. Cane sugar is used because of its sweetness, while milk sugar is not nearly so sweet. When several compounds have the same formula, they are said to be isomers. The best explanation for such a happening, which occurs only among the organic compounds, is that the extreme complexity of their molecules makes possible different arrangements of the atoms.

In other words, the arrangement of the carbon, hydrogen and oxygen atoms differs in all three sugars even though the number of the atoms is identical. To illustrate this, simpler substances than the sugars may be used. There are three hydrocarbons of the paraffin series known by analysis to have the formula C_5H_{12}. Two of them, pentane and isopentane, occur in petroleum. The third, tetramethyl methane, is prepared in another way. To help explain this state of affairs the formula can be considered in three different ways.

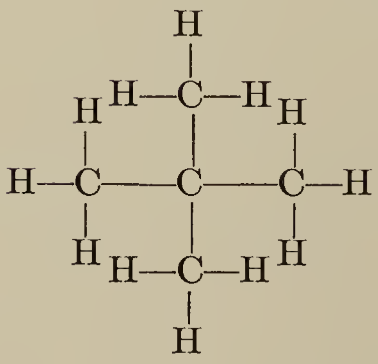


$$\begin{array}{ccccccccccc} & & H & & H & & H & & H & & H & & \\ & & | & & | & & | & & | & & | & & \\ H & — & C & — & C & — & C & — & C & — & C & — & H \\ & & | & & | & & | & & | & & | & & \\ & & H & & H & & H & & H & & H & & \end{array}$$

pentane or



isopentane or



tetramethyl methane.

Formulas written this way are called graphic or structural formulas. While they do not represent the exact structure they do represent the idea of a definite arrangement of the atoms in the molecule. They also represent carbon as having a valency of four, and hydrogen of one.

The molecule of pentane represented above would seem to have only two dimensions while probably the molecule has three, and one must think of it in space with its atoms arranged in definite order much as one imagines the figures in solid geometry. Sometimes it is possible to make structural formulas of several isomeric compounds while only one compound has as yet been isolated by the chemist. But there are never any more compounds actually known than the possible number indicated by structural formulæ, a fact which substantiates our theory that the differences are due to arrangements of atoms. Pentane and its isomers are comparatively simple, but the sugars with a constitution represented by $C_{12}H_{22}O_{11}$ offers the possibility of a greater number of structural formulæ, so that three isomeric sugars are not so strange.

The close relationship between the inorganic and the organic compounds, and yet the many peculiar characteristics of the latter, makes some elementary knowledge of them necessary as an introduction to the subjects of foods and cleaning substances; for most of these substances are organic compounds, or compounds of carbon.

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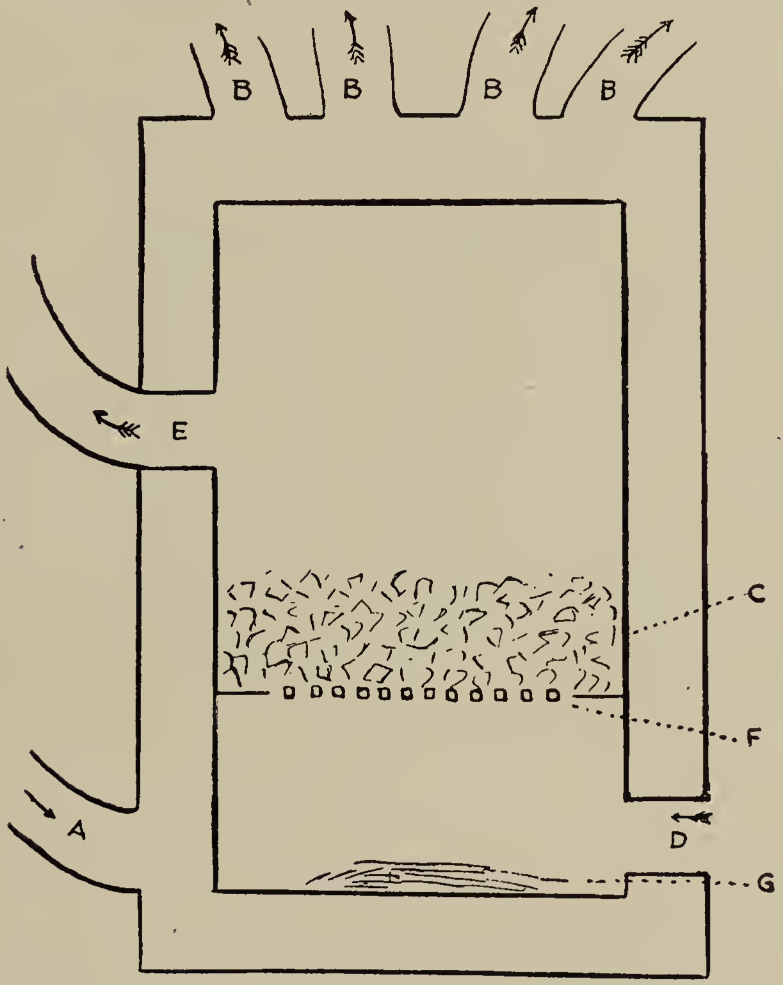
CHAPTER II

FOOD IN THE BODY

THE part played by food in the body can best be understood by comparing the body to a furnace. A furnace is a machine built by man in such a way that the chemical energy in fuel, such as wood or coal, may be liberated by its uniting with oxygen, or burning. The furnace is made of iron to resist the heat, and is so planned that the air enters below the fuel, mixes and combines chemically with it, the gases formed pass up the chimney and are disposed of and the heat is used to raise the temperature of the air in the house either directly or by means of steam or hot water. If the furnace wears out in parts these must be repaired from without. The ashes too must be removed periodically or else the fire will not burn.

Now the body of a man, or of any animal in fact, bears many striking resemblances to this furnace. The body also is a machine in which the energy in substances used as foods, such as meat and bread and vegetables, may be set free, here too by their union with oxygen, and this energy used for the purposes which the body requires. The body too needs its supply of oxygen, which is breathed into the lungs, so that the foods may be oxidized. The waste products resulting from this oxidation must, like the gases and ashes of the furnace, be got rid of or the body cannot live in health. But here the similarity ends. Instead of being built of some material which can resist the oxidation of an inner fire, it is really the substance of the body itself that is burned, and instead of requiring outer assistance to repair its worn or broken parts, it has marvelous powers of doing its own repairing from within and, more than that, of increasing in size, or growing. In other words, it is a living organism instead of a lifeless machine.

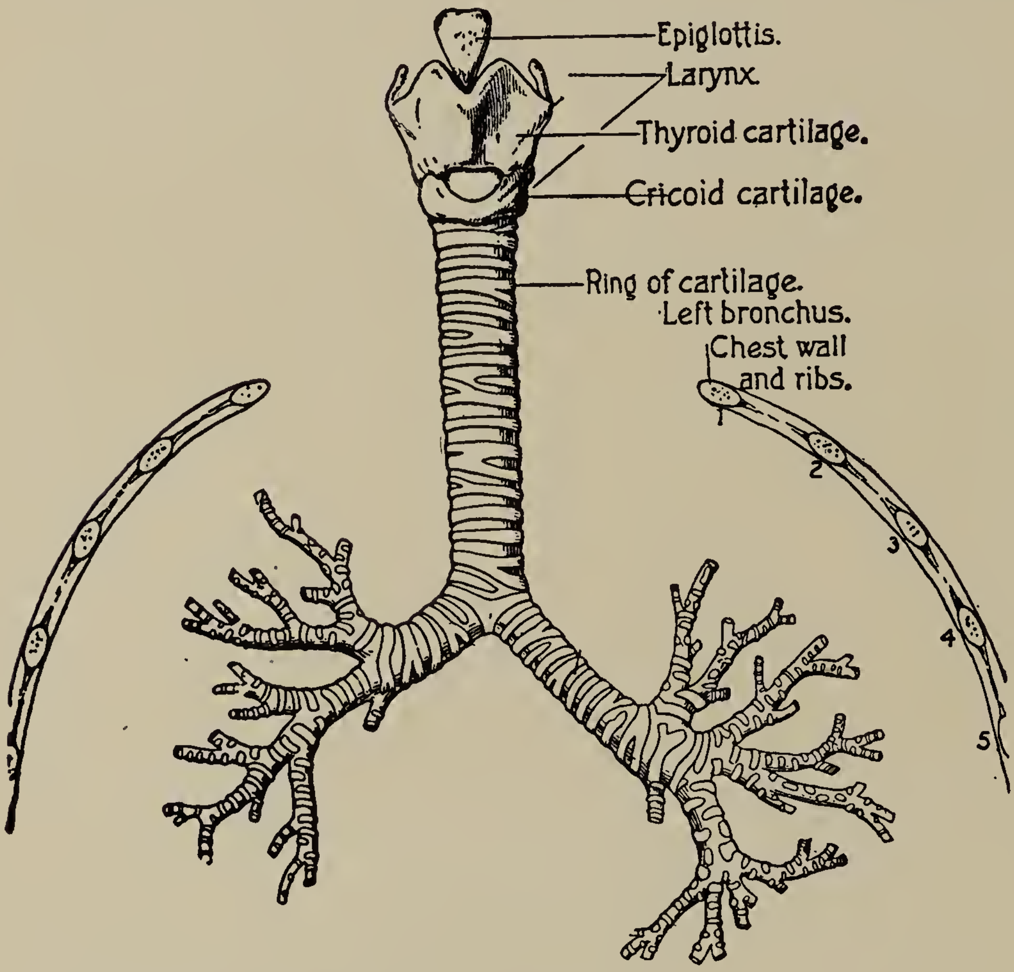
It will be remembered that the air is four-fifths inert nitrogen and about one-fifth oxygen. When the chest is expanded and air rushes into the lungs this oxygen is passed by diffusion through the walls of the air sacs in the lungs into the blood vessels, where it combines chemically with a substance in the blood called hæmoglobin, and forms oxyhæmoglobin. So fresh blood containing



A Hot-Air Furnace

a. Incoming fresh air. *b.* Warm air pipes to rooms. *c.* Bed of coal. *d.* Incoming air for fire. *e.* Outgoing smoke and foul air to chimney. *f.* Grate. *g.* Ashes.

oxygen is constantly flowing away from the lungs and carrying the oxygen in this loose combination with haemoglobin through all parts of the body. When food is taken into the digestive tract it undergoes processes called by the general term "digestion," which are largely chemical in their nature, and in a soluble form the products of digestion are passed through the thin lining of the tract into the blood flowing constantly away to all other parts of the body. In any active tissue of the body, as for instance the muscular tissue of the hand, these food substances are taken up from the blood by the cells, there

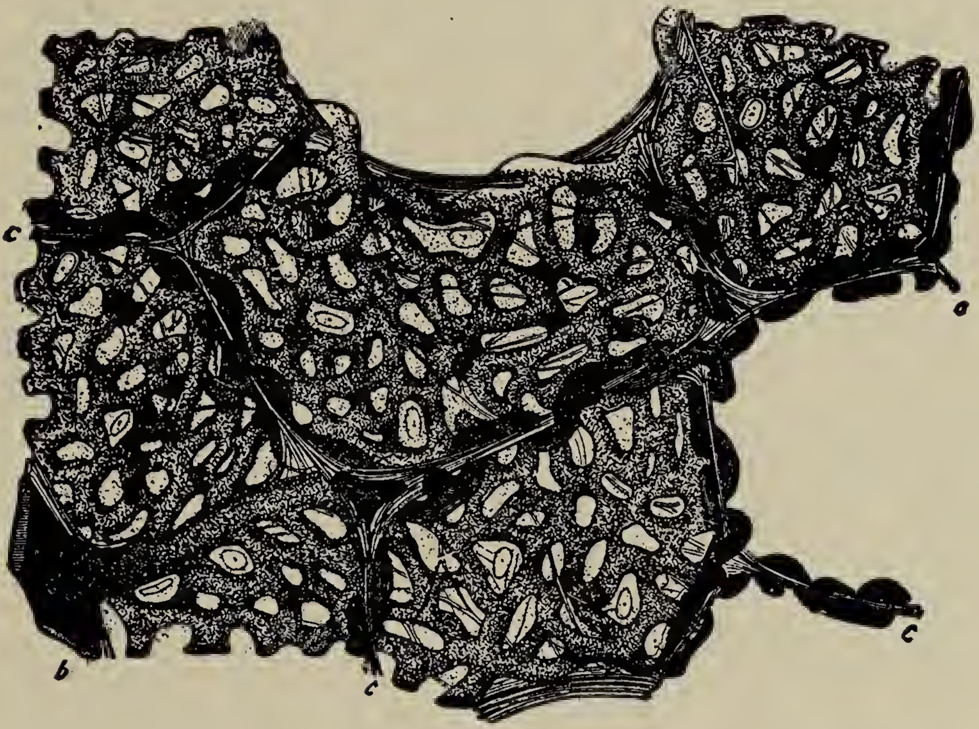


From "Principles of Physiology and Hygiene"
By G. W. Fitz, M.D.; published by Henry Holt and Company

Front View of the Trachea and its Branches

In the air sacs at the ends of the smallest divisions of these tubes the oxygen passes through the walls and into the blood in the capillaries.

undergo further chemical changes and so become actually part of the muscular tissue, and are ready, when the time for work comes, to combine, almost explosively we may say, so sudden is the change, with the oxygen also brought by the blood, and so to set free the chemical energy of these substances. This energy is used partly to keep the body warm, its normal temperature being 98.4° F., and



From "Principles of Physiology and Hygiene"
By G. W. Fitz, M.D.; published by Henry Holt and Company

Section of Lung with Distended Blood Vessels, Highly Magnified

c. c. Partitions between air sacs, at the ends of the smallest bronchial tubes.
b. Small artery giving off capillaries to walls of air sacs.

partly to do the mechanical work of the body, such as muscular contraction, secreting of juices and so on.

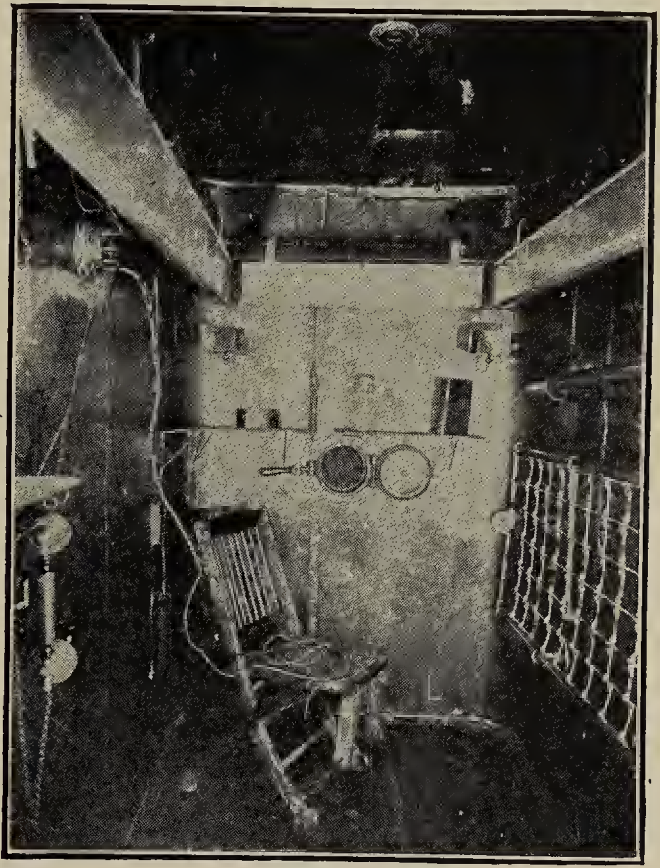
One of the most remarkable features of these changes, and most difficult for us to understand, is the nature of this oxidation. Oxidation of such substances as make up food, if taking place outside of the body, requires a much higher temperature than that of 98.4° F., which is maintained in the tissues. Another remarkable circumstance is that these bodily oxidations take place in very moist substances, which would certainly prevent rapid chemical action under other circumstances. There is some control of these processes in living tissues, some difference due to the fact that the body is not merely so much lifeless matter but controls its own action, which is, as yet at any rate, beyond our understanding. The facts, as far as results go, are beyond dispute. An explanation of their circumstances has yet to be formed.

From U. S. Department of Agriculture Office of Experiment Stations, Bulletin 175

The laboratory containing the respiration calorimeter — the small room in the background with the opened window — and appliances for supplying fresh air, removing used air, and measuring the heat given out by the occupant of the room under different conditions of nutrition and activity.

The waste products resulting from these oxidations must be got rid of, so that they may not clog the working parts of the body, and so hinder their action, just as the gases and ashes must be removed from the furnace. The blood here again serves to carry these products away to the organs of excretion, the lungs, the skin, and the kidneys, and by them these wastes are removed.

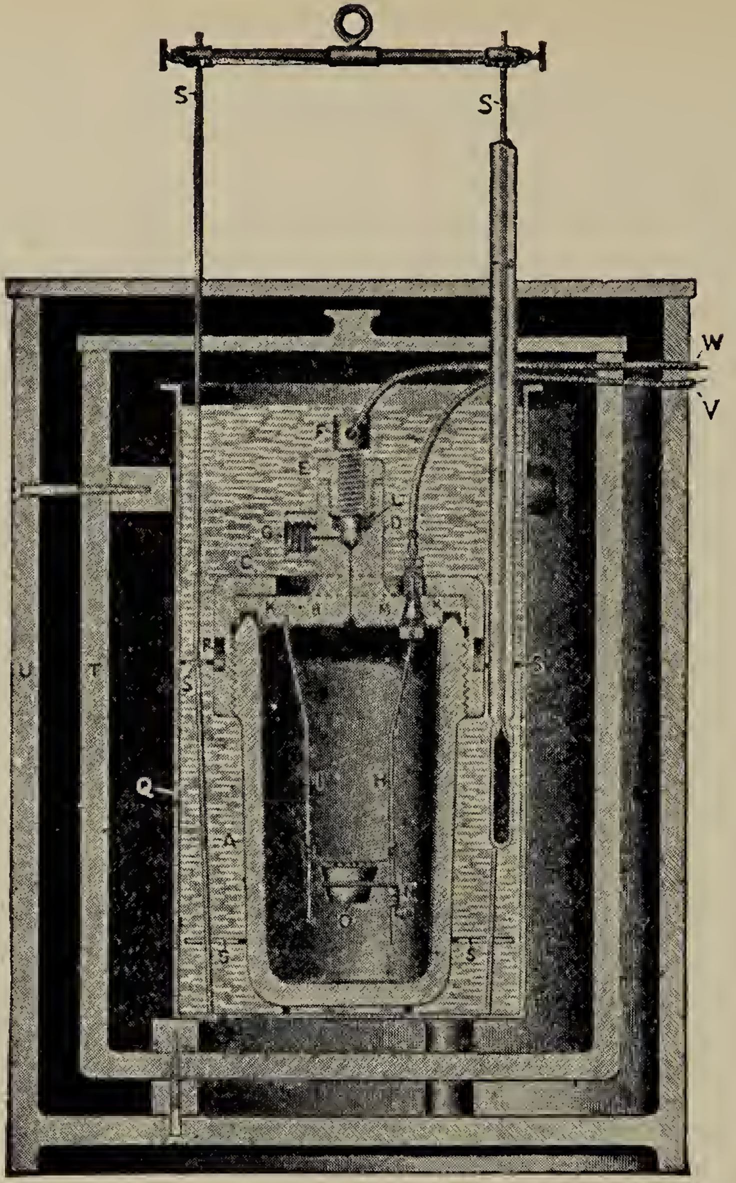
The uses of food in the body may be looked upon as twofold, although these uses are not distinct from each other. First, food is needed to build up new tissue in the growing body, and to repair worn-out parts. Second, food is needed so that by its oxidation the body may make use of the energy contained in it to keep the temperature normal and to do the work of the various organs. The value of any food must depend on its ability to perform one or the other of these functions. Later we shall classify foods, and it will then be seen that some classes are used specially to build up tissue, while others are pre-eminently energy-producing foods. The amount of energy that any food can give up for the use of the body is determined by drying it and then burning it in a calorimeter, or instrument for determining its heat value.



From U. S. Dept. of Agriculture Office of Experiment Stations, Bulletin 175

An Interior View of the Respiration Calorimeter

Very recently a class of substances called vitamines has been discovered, which adds a new group to the classification given in the following chapter. These vitamines are present in minute quantities in fresh foods, such as fruits, vegetables, milk and butter. Their chemical composition is unknown, but their influence on the well-being of the body is important. They do not serve to



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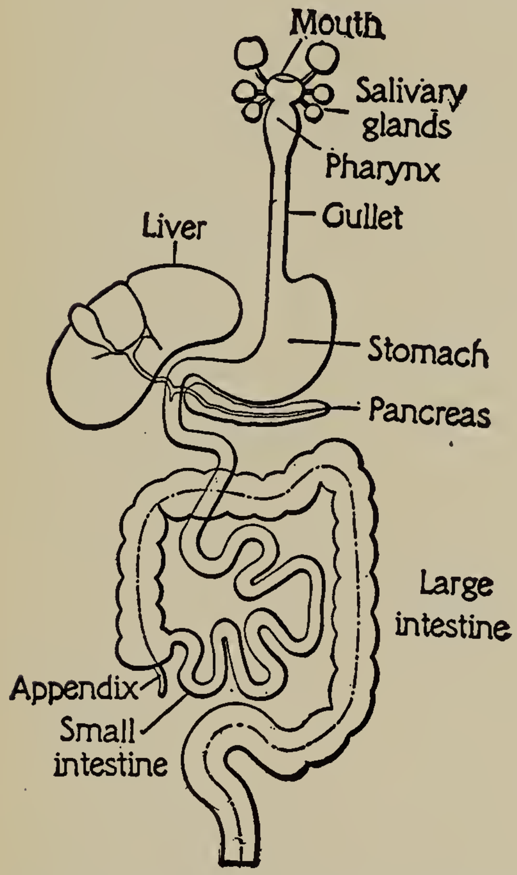
Bomb Calorimeter

U. and *T.* fiber pails with air space between.
Q. metal cylinder containing water whose change of temperature is to be observed.
S. S. stirrers for equalizing the temperature of the water. The thermometer is carried by these.
A. cylindrical cup containing oxygen.
O. platinum capsule holding substance to be tested.
I. H. platinum wires holding the capsule and connected with
W. V. rubber-covered wires conducting the current.

A fine iron wire, coiled into a spiral over the substance to be burned, connects *I.* and *H.* The current passes through this, burns it, and it drops upon the substance in the capsule, so igniting it. The fuel value of this substance is determined by the change of temperature in a known weight of water in *Q.*

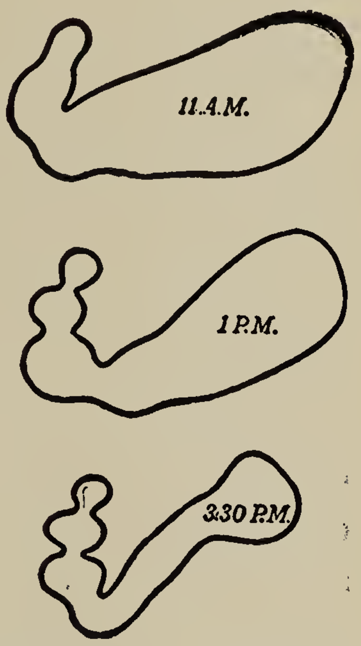
For a description of the bomb calorimeter and its use see pp. 659–699 in the journal referred to.

produce living tissue, nor to furnish heat and energy, but they seem to act as regulators of bodily activity, and by their presence to set free energy for the life processes of the body. Cooking impairs their action, wholly or in part, and so the necessity for eating regularly some raw food is evident. We shall probably hear more from investigators about vitamines in the future.



From "Principles of Physiology and Hygiene" By G. W. Fitz, M.D.; published by Henry Holt and Company

Diagram of the alimentary tube and its appendages. (Testut.)



From "Principles of Physiology and Hygiene" By G. W. Fitz, M.D.; published by Henry Holt and Company

Diagram illustrating the muscular contractions of a cat's stomach after eating a meal just before 11 A.M. By these contractions the contents are squeezed toward the pylorus.

In order that food may thus serve to build up the tissues, and to produce heat and energy by its oxidation, it must usually undergo changes in the digestive tract. These changes begin in the mouth, where food is ground and crushed by the teeth, softened by the saliva, some soluble foods dissolved, and some acted on chemically by a substance in the saliva. This work begun in the mouth is continued in the stomach, which is primarily a receptacle for retaining the food until it is sufficiently softened and separated to pass into the intestines. The most important chemical changes take place in the intestines, where several juices complete the work of digestion, and where most of the food now ready to be assimilated or made over into the substance of the tissues is absorbed into the blood. The undigested part of the food taken is finally passed out of the large intestines as waste. The chemical changes undergone by the various kinds of foods are complicated, and are still the subject of much investigation, but in general it is believed that they are processes of hydrolysis followed by a splitting up of large molecules into smaller ones, a process which seems necessary in order that they may be absorbed. The chemical changes undergone by the food depend on their chemical composition, and this can best be understood by studying the main classes of nutrients.

QUESTIONS

1. What is done to the body that may be compared to stoking a furnace? To opening the draft at the bottom?
2. To what use are put both food in the body and coal in a furnace? To what use is food put that coal cannot serve?
3. What changes not of a strictly chemical nature does food undergo during digestion?
4. Why does food need to be made soluble during digestion?
5. When two bottles of different gases, air and hydrogen for instance, are placed mouth to mouth for a few minutes, the gases are found to be equally mixed. What is this phenomenon called? Does it occur with substances in solution? Explain.
6. In what three respects are the rusting of iron and the use of food materials in the body similar?

7. What source of error might be furnished by determining the fuel value of food by burning it in a calorimeter? Would the estimate probably be too high or too low?

8. Are waste products of bodily activity — other than those excreted by the intestines — necessarily soluble? Give reason for your answer.

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CHAPTER III

WATER AS FOOD

During the study of carbon compounds, some of these substances, such as alcohol, wood, sugar, were burned, and it was discovered that they all contained carbon, oxygen and hydrogen. Horn, which is really the substance of which the nails and hard outer skin of animals is composed, was found to contain nitrogen also. Compounds of these four elements make up the chief part of the food for all living things, both plants and animals. These four elements, together with small quantities of metals, such as sodium, potassium, calcium, iron and magnesium, and a few non-metals, such as sulphur, phosphorus, fluorine, chlorine, bromine and iodine are the only ones of the eighty or ninety known elements that are found in the protoplasm of living organisms, or in the substances that make up their food.

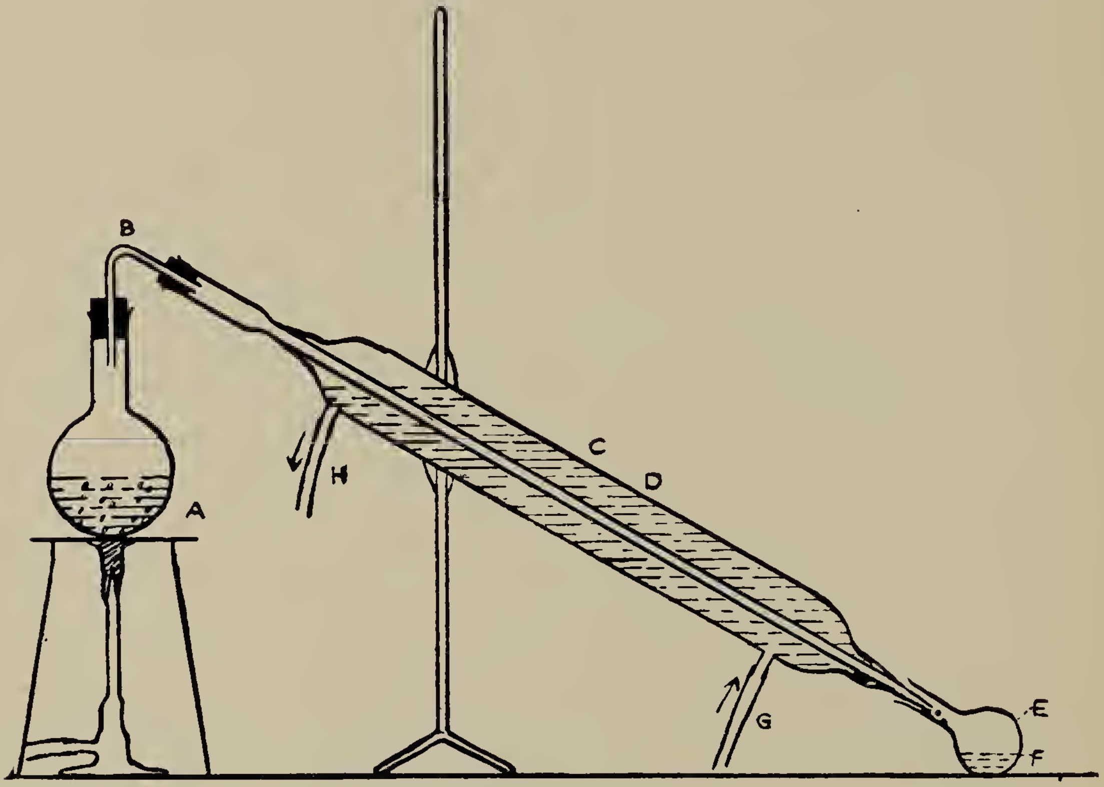
It would be an endless task to study foods singly. Any simple meal is made up of a number of foods, — bread, butter, milk and meat being only a few of the numerous different foods used by man. However, when various foods are analyzed, it is found that the many substances which they contain may be grouped into a few classes. These we call food-stuffs or nutrients. This classification makes the study of foods a much simpler matter, since in each class all the substances are alike in composition, need similar treatment in cooking, are subjected to similar processes of digestion, and perform similar functions in nourishing the body and providing heat and energy.

These nutrients may be separated into two large groups, those undergoing oxidation in the body, and those not oxidized. Under the former heading come carbohydrates, fats and proteins; and under the latter heading come inorganic salts and water. We shall begin the study of nutrients with water.

The chemical composition of pure water has already been studied. It has been electrolyzed into hydrogen and oxygen, two volumes of hydrogen to one of oxygen. But water is so good a solvent that chemically pure water is never found in nature and never used in our households. The nearest approach to it probably is found in rain water, but this contains gases dissolved from the air, dust and bacteria. Water that has fallen on the earth and run over or through the ground is likely to contain in addition appreciable quantities of mineral salts which it has dissolved out of the soil. These minerals in solution give their special characteristics to mineral spring waters such as the famous bath and medicinal waters. They also cause the hardness often found in the waters of certain localities and which is so objectionable when the water is used in boilers, or for laundry purposes. This hardness of water will be referred to again under the work on cleaning. For drinking purposes small amounts of these salts are harmless and even desirable to relieve the otherwise flat taste.

Chemically pure water may best be obtained, if needed, by distillation. A colored solid like copper sulphate may be dissolved in some water, so that its presence is evident. This water is boiled in a flask whose only outlet is through the inner tube of a Liebig's condenser, where the steam is surrounded by cold water flowing through an outer jacket, and so condensed and collected in a receiving flask. This distillate is found to be perfectly colorless water with no trace of any solid impurity. All substances with higher boiling points than that of water are left in the first flask.

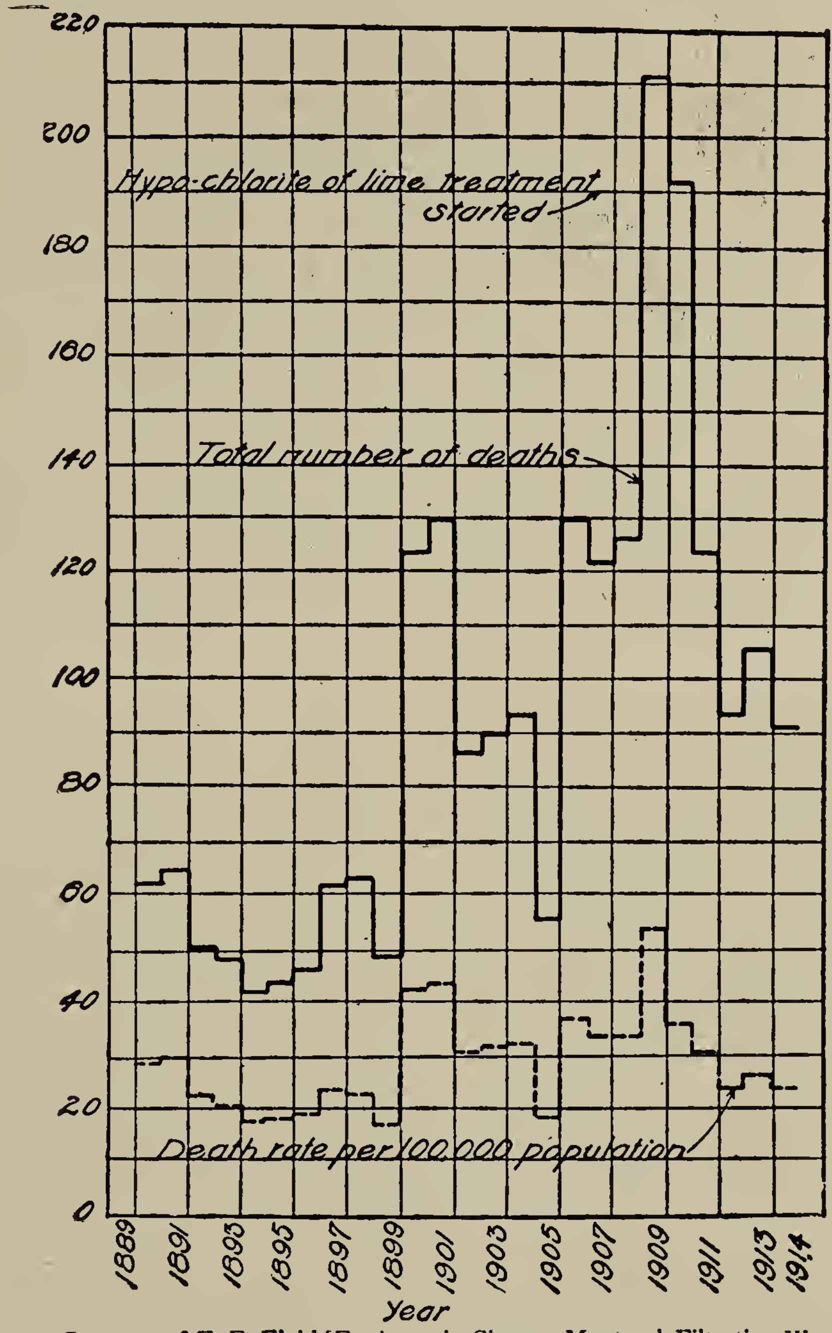
Water for drinking purposes fulfills all requirements if it be of an agreeable taste, clear in appearance, and free from pollution by sewage, for this means the probable presence of harmful bacteria. Such a water may best be obtained, not by distillation, but by a process of filtration. On a large scale for city use there are filter beds or large areas of carefully graded gravel and sand to which the bacteria cling as the water trickles through. A layer of slimy material is gradually formed on top which is effectual in holding bacteria but which must be removed periodically. Sometimes a gelatinous precipitate of aluminum hydroxide is formed in the water by adding



Liebig's Condenser

a. Flask containing water to be boiled. The steam passes out through *b* and into the inner tube *c*. Here it is condensed by cold water flowing through an outer jacket *d*, and drops into a receiving flask *e*. *f*. Condensed water. *g*. Inlet for cold water. *h*. Outlet.

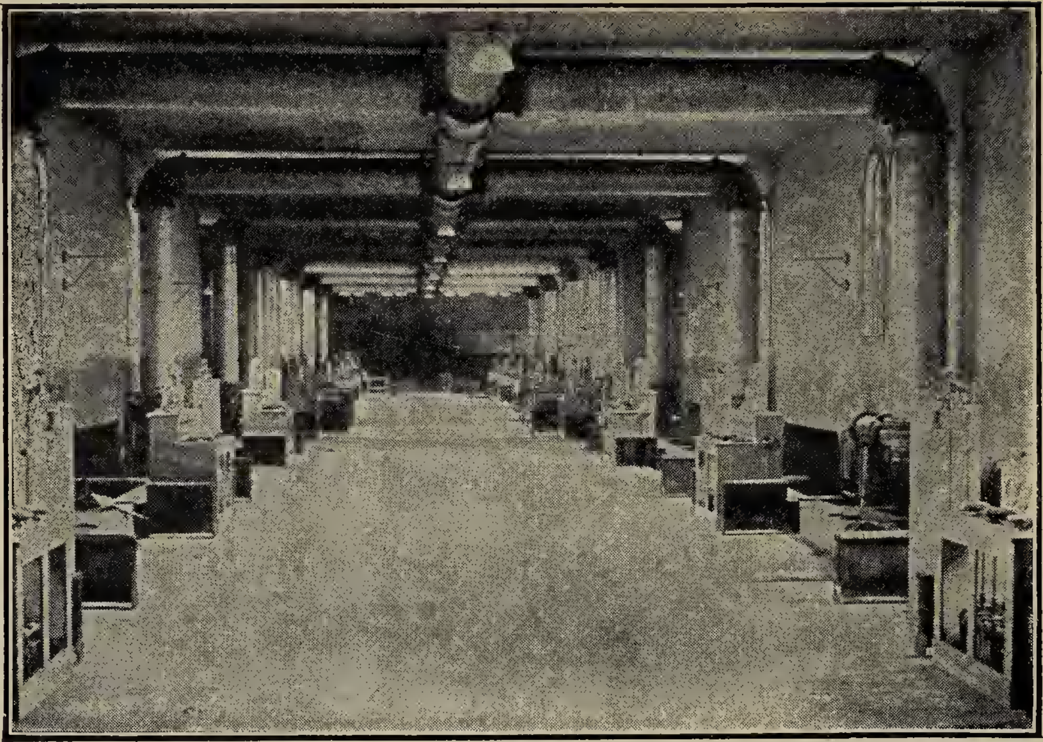
alum (or aluminum sulphate) and sodium carbonate, and this forms the slimy layer more quickly and furnishes quicker filtration. Such water is sometimes made doubly secure against infection by the addition of a very small



Courtesy of F. E. Field, Engineer in Charge, Montreal Filtration Works

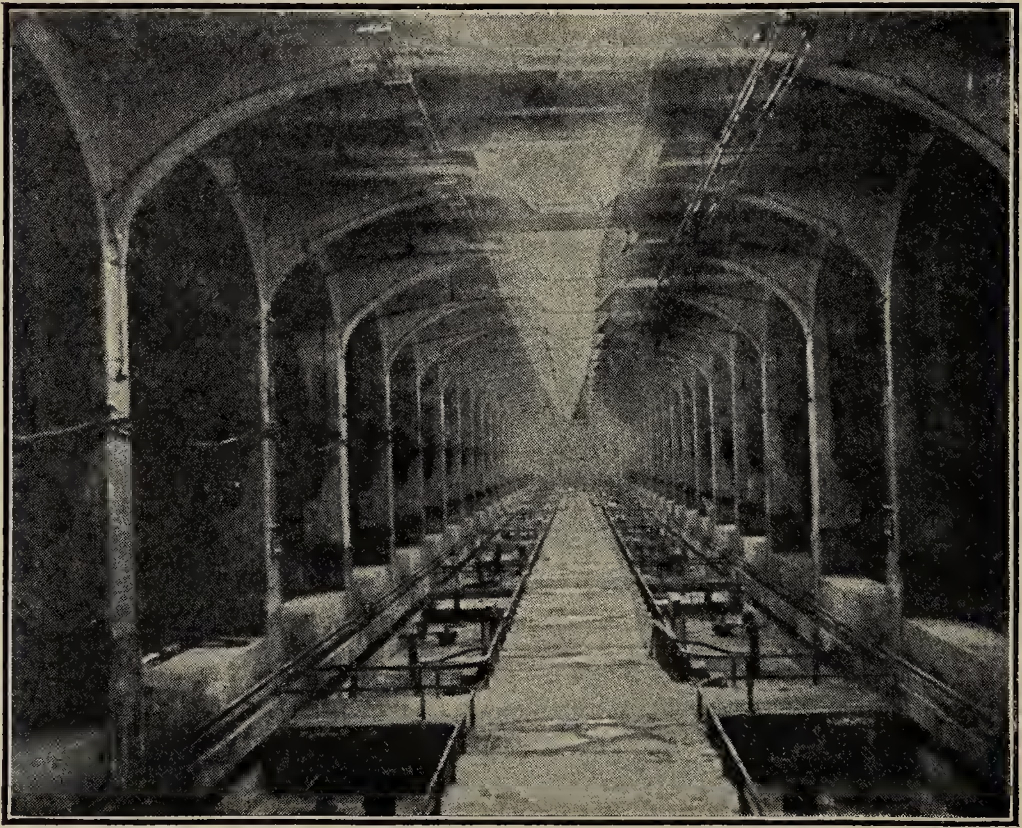
Chart showing Deaths from Typhoid Fever in Montreal, 1890–1914

This shows the necessity for some public means of purifying the water of large towns and cities. The treatment with hypochlorite was begun before the present filtration system was undertaken.



Courtesy of F. E. Field, Engineer in Charge, Montreal Filtration Works

Prefilter Gallery



Courtesy of F. E. Field, Engineer in Charge, Montreal Filtration Works

Final Filter Gallery

The beds of filtering material are at the sides. Electrical apparatus for controlling the water flow is seen in the galleries. There are elaborate systems for cleaning the sand, and for automatically adjusting the amount of water leaving the prefilters and filters to the amount pumped from the reservoir.

amount of sodium or calcium hypochlorite or a similar chlorine compound which destroys any bacteria remaining.

EXPERIMENT 15

Purpose.

To clear water of sediment with alum.

Directions.

Mix some fine clay with water, and then dissolve some alum (or aluminum sulphate) in it. Add a little concentrated sodium carbonate solution, and stir. Result? Allow to stand several hours. Result?

Questions. Write the word equation. Water enters into the reaction. Name the solid substance. What is its condition as to solubility?

In the household the water furnished from well or spring in the country, or from the reservoirs in the city, is usually free from suspended sediment. There may be times when it is desirable to clear it or to be especially sure that it is free from bacteria. The methods then used are usually filtration through artificial stone, porcelain or charcoal, and boiling. Stone or charcoal filters render the water clear, but they do not guarantee it against bacteria. Unless the charcoal be thoroughly aired and cleaned very frequently, such a filter is too often likely to be a breeding place for bacteria which may finally pass on with the water leaving the filter. The best method of sterilizing water in the household is to bring it to the boiling point. This kills all active disease germs. To insure against the resting spore state, the water should be brought to the boiling point, then cooled, and finally boiled ten or fifteen minutes.

It is a common mistake to think that freezing purifies water and that therefore ice from any source may be used for drinking purposes with impunity. As a matter of fact, there are bacteria capable of withstanding the freezing process, and when the ice is thawed they are ready again to become active. The source of the ice supply should therefore be carefully inspected whenever the ice is used in food or drink.

In testing water for drinking purposes to discover whether it contains bacteria that might cause diseases such as typhoid fever, the effort is not usually made to determine the presence of the bacteria themselves, but to find out whether there is present nitrogenous organic matter of animal origin, in other words, sewage contamination. If this be present there is every reason to suppose that harmful bacteria accompany it. Such organic matter is unstable, that is, is easily oxidized, and its presence may be detected by such an experiment as follows in part *b*.

EXPERIMENT 16

Purpose.

To study simple tests on drinking water.

a. To show the solvent power of water.

Directions.

Filter some faucet water if necessary so that any suspended particles are removed. Evaporate to dryness in an evaporating dish over the flame, being careful not to overheat the dish when dry. What remains? How was this held in the water? What kind of substances make it up?

Directions for parts b, c, d and e.

For each part use three test-tubes labeled 1, 2 and 3. In each case half fill 1 with distilled water, 2 with tap water, and 3 with distilled water to which has been added the substance whose presence is to be tested. The purpose is to observe whether 2 reacts like 1 or like 3.

b. To show unstable organic matter.

Add to each tube a few drops of concentrated sulphuric acid, and just enough potassium permanganate solution to make the liquid a pale pink. Add to test-tube 3 a little sewage. Results?

NOTE. The permanganate is an oxidizing agent, and if organic matter that may be oxidized is present, the permanganate is decomposed and its color disappears.

c. To show chlorides.

Add to test-tube 3 a little of a solution of any soluble chloride. Add to each tube a few drops of nitric acid and then a little silver nitrate solution. Do not heat. Results? The nitric acid simply prevents the precipitation of carbonates if present.

d. To show sulphates.

Add to test-tube 3 a little of a solution of any soluble sulphate. Add to each tube a few drops of hydrochloric acid, and then a little barium chloride solution. Results? The acid again prevents the possible precipitation of carbonates.

e. To show calcium compounds.

Add to test-tube 3 a little calcium hydroxide solution called lime water. Add to each tube a few drops of ammonium oxalate solution $(NH_4)_2C_2O_4$. Results?

Questions. Does your faucet water respond to any of these tests? Which ones?

Write an equation in symbols for the reactions in test-tubes 3 of parts *c*, *d* and *e*.

NOTE. Parts *c*, *d* and *e* are tests for classes of inorganic or mineral salts soon to be studied as a class of nutrients. Calcium salts are a frequent cause of hard water.

Mineral salts are a disadvantage in water used for the laundry, but they are not harmful in drinking water. However, in testing water for drinking purposes, tests are made for certain nitrogen salts, as ammonium salts, nitrites and nitrates, and the relation of these to the amount of chlorides present is an indication to the trained chemist of the probability of sewage contamination, since these salts are either excreted directly by the body or are formed by oxidation of its excretions.

The importance of water to the body cannot be overemphasized. The evaporation of water excreted from the skin in perspiration is the most important means of regulating the temperature of the body. Not only does water constitute a large percentage of the weight of the blood and lymph and of various secretions or liquids manufactured by glands, and then poured out to do some work in the body; but the working tissues themselves, such as muscle fibers, nerve cells and so on, made up of living protoplasm contain a large proportion of water. Beneath the hard protective covering of the body the soft underlying parts are composed of cells whose contents are semi-fluid, and these are bathed in lymph, which is largely water. It has been calculated that about seventy-five per cent of the weight of the human body is due to water. No food with the exception of oils, can be absorbed into the blood without being at least dissolved in water. It is evident therefore that the purity of the water supply should receive much consideration.

EXPERIMENT 17

Purpose.

To show the presence of water in organic substances.

Directions.

Heat in test-tubes separately a little of the following substances, holding the open end of the test-tube slightly lower than the closed end.

Meat, wood, carrot, cabbage, parsnip, potato, peas, or any fresh or dried vegetable obtainable. What evidence is there that these substances contain water? Refer to the table of percentage compositions of different foods, and note the percents of water present.

QUESTIONS

1. Could water be freed from ammonia by distillation? From hydrogen sulphide? From sodium chloride? Give reasons for your answers.

2. Is a clear, sparkling water always safe to drink? Give reasons for your answer.

3. What are the dangers from well water?

4. Why are chlorides in water an indication of danger?

5. Which has a larger percentage of water, milk or cabbage? (See table of Food Compositions, Appendix C.)

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CHAPTER IV

SALTS AS FOOD

EXPERIMENT 18

Purpose.

To show the presence of mineral salts in bone and in foods.

Directions.

a. Get a cooked bone, and chop it into pieces about a cubic inch in size. Place a piece in dilute hydrochloric acid in a small bottle, and stopper loosely. Describe the action. Allow it to remain twenty-four hours or longer. Test the gas in the bottle by all the tests you know. What is it? Filter and evaporate to dryness some of the liquid. Result? Test for calcium as in Experiment 16 *e.* Examine the bone and describe it.

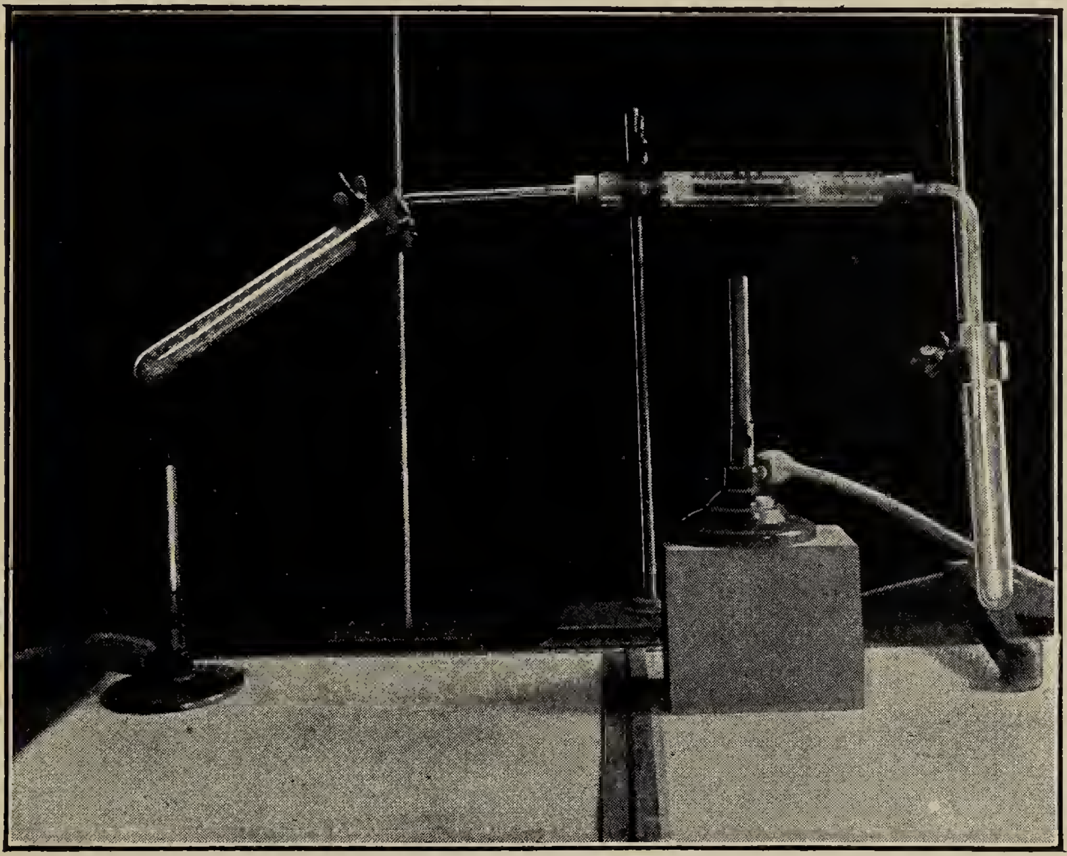
> **Questions.** What kind of substance is left from evaporation? What kind of salt in the bone must have caused the formation of the gas? Write in symbols a possible equation.

b. Each pupil may use one or several foods. Place some oatmeal, beans, meat, cheese, milk, cocoa, bread, celery, asparagus, fruit of any kind, in small pieces on an iron pan or in a porcelain crucible, and heat over the hottest flame for one hour, or until no further change occurs. Describe what happens, and the appearance of the residue.

The residues from all the different substances may conveniently be heated together at the lecture table. Place them in a hard-glass tube connected at one end with an oxygen generator, and arranged so that a delivery tube from the other end opens into limewater. Heat the tube gently, and pass oxygen through it slowly. Describe the action and the appearance of what remains when action ceases. Moisten this residue and test it with litmus paper. Is it acid or alkaline?

Questions. What was the black substance? What became of it? Why does the residue not disappear likewise? Are mineral substances oxidizable?

The mineral part of bone is made up chiefly of calcium and magnesium carbonate, and calcium and magnesium phosphate. Both are dissolved by the hydrochloric acid in the experiment, and the carbonates react with it to give calcium or magnesium chloride, carbon dioxide and water. When the residue from the burned foods is heated in the combustion tube, the carbon not already burned is now burned, and the salts are decomposed by the heat, leaving, if the temperature be high enough, only the oxides. In the body this, of course, does not happen. There the salts may be excreted as they are received, as



Residue from burned food ready to be heated in a stream of oxygen.

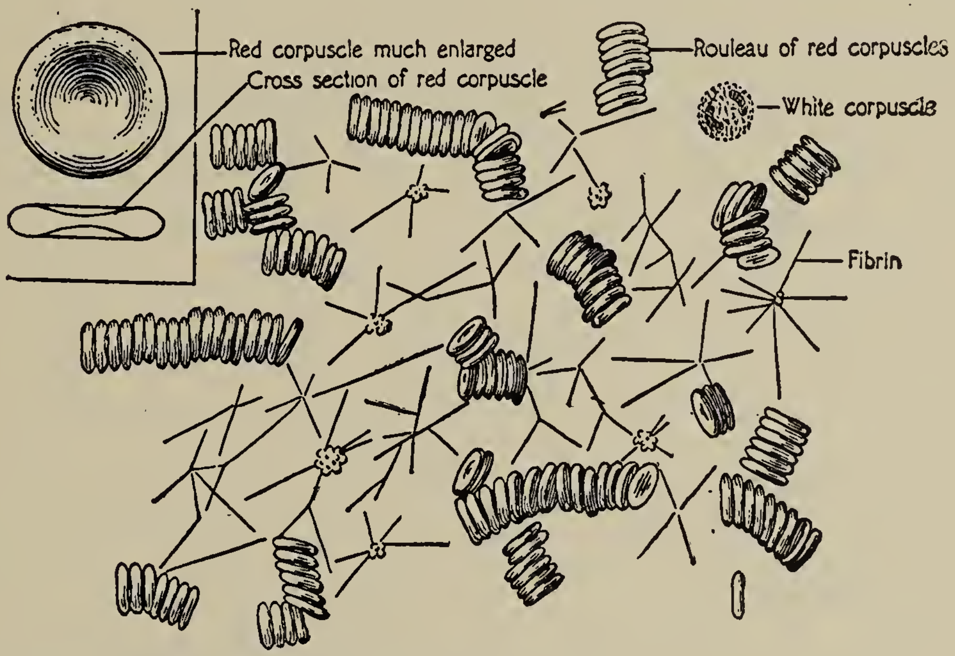
is the case with sodium chloride, or they may be formed by the metabolism of the body, that is the chemical processes that take place during nutrition, and be excreted from the kidneys or intestine, as with sulphates and phosphates.

It was shown in the study of water as a food that inorganic salts may be taken into the body dissolved in drinking water, which has acquired them in its passage through the soil. The experiment just performed shows that such salts are present in our own bodies, especially in the hard tissues such as bone and teeth; and that we take them into the body in many different foods. The table of food values shows the relative proportions of such salts in common foods. The ashes left from a wood or coal fire consist of these non-oxidizable, inorganic salts. From the chemical point of view those found in our food consist chiefly of the phosphates, sulphates and chlorides of calcium, potassium, sodium, iron and magnesium.

Their percent by weight in foods is small, but their value to the body is out of all proportion to their quantity. In addition to their use in the bony structures, where they make up the rigid and resisting frame-work on which the softer, active tissues depend for support and protection, they are essential in the protoplasm of the active tissues, and are also necessary in the digestive tract and in the blood.

The mineral salts are the only class of nutrients containing to any extent metals or base-forming elements. They are found chiefly in ripe fruit, grasses, fresh green vegetables and milk. Animal foods generally and cereals supply the larger part of the acid-forming elements. Now foods of animal origin, especially proteins, during digestion in the alimentary canal, are particularly liable to undergo decomposition due to the action of bacteria, and to produce substances poisonous to the body if absorbed. Some of these decomposition products are acids, and it is plain that to neutralize them the base forming elements are necessary. This shows how valuable vegetable foods are in any diet, and why the lack of such foods may lead to many disorders.

Another view of the subject may be had from the condition of the blood. Recent investigations have proved that mineral salts are necessary here. Dissolved in the liquid part of the blood, the plasma, they increase its weight and so regulate its specific gravity that the small bodies, the corpuscles, which float in it are able to retain a normal size and shape. A diet containing too little mineral food is very soon followed by a plasma of less density, and under such conditions water from the blood passes through the walls of the corpuscles and they become distended and misshapen. Many important activities

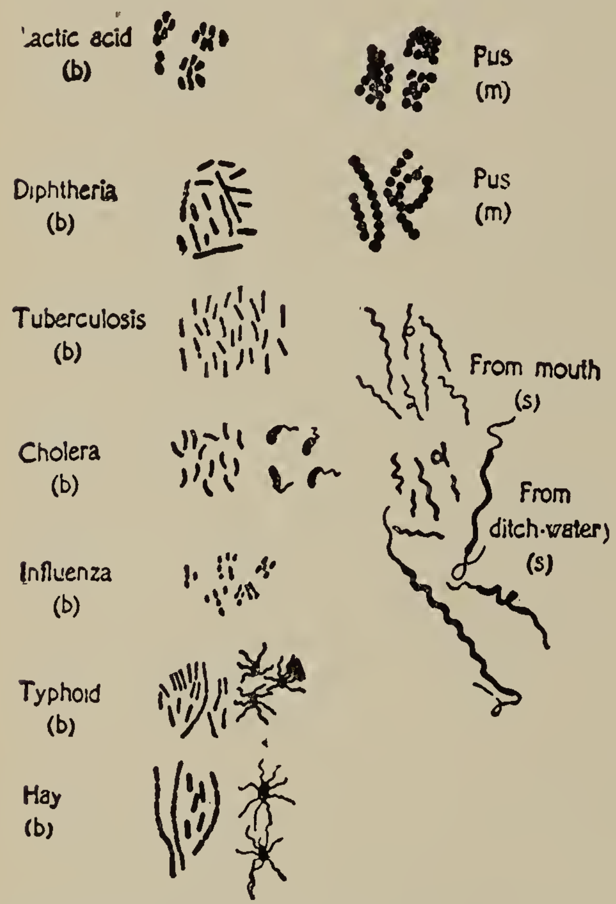

Red corpuscle much enlarged
Cross section of red corpuscle
Rouleau of red corpuscles
White corpuscle
Fibrin

From "Principles of Physiology and Hygiene"
By G. W. Fitz, M.D.; published by Henry Holt and Company

Coagulated or Clotted Blood, Highly Magnified

In the body the red corpuscles do not form such rolls, and the fibrin has not separated out from the plasma, or liquid of the blood.

of the body are dependent on these corpuscles, and any lessening of their working power is followed by serious results. It may be that immunity to such diseases as tuberculosis, typhoid fever, scarlet fever and diphtheria is dependent on such a proper proportion of minerals in the food that the specific gravity of the blood and so the shape and functioning of its corpuscles may be normally maintained.



From "Principles of Physiology and Hygiene" By G. W. Fitz, M.D.; published by Henry Holt and Company

Various forms of bacteria, highly magnified, some of which have flagella or cilia. *b.* bacilli. *m.* micrococci. *s.* spirillae.

Vegetable foods such as cereals, to furnish their proper proportion of inorganic substances, should not be subjected to too much refining. It is possible by refining processes to obtain foods of high calorific value, capable of building up the tissues and furnishing energy, which nevertheless lack the mineral salts. For instance the calcium salts found in the outer coats of grains are valuable in building the teeth and bones of growing children. This is one reason why whole wheat flour is preferable to white flour for bread, and why unpolished rice is more wholesome than the polished grain. The refining of barley is a case in point. Barley is subjected to at least six pearling or refining processes, which remove in succession more and more nutritious material, so that when put on the market it has lost about seventy per cent of its nitrogen, eighty-five per cent of its K_2O, ninety-two per cent of its P_2O_5. If the public were willing to accept it after the second pearling, when all the husk has been removed, millions of pounds of nourishing food would be saved.

The exact part played by mineral foods in the body, especially in the blood and internal secretions, offers a wide field for investigation and experiment on the part of chemists, in which much may be discovered of great importance to the physician and dietitian in their work of restoring healthy conditions in the body and maintaining them.

QUESTION

Why are young children often given limewater in milk?

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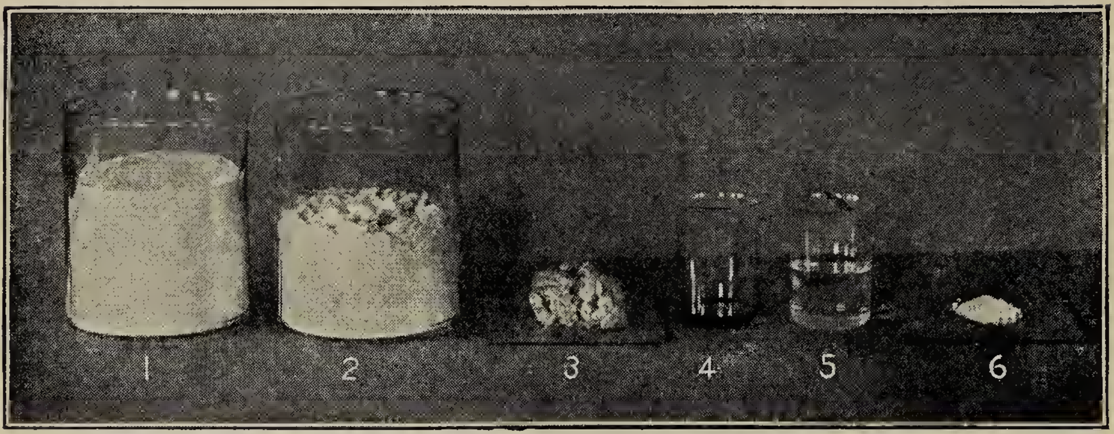
CHAPTER V

OXIDIZABLE NUTRIENTS



Analysis by Students.

Analysis of One Pound of Flour

1. 1 lb. (3 cups) flour.
2. 14 oz. carbohydrates in 1 lb flour.
3. 1.7 oz. proten in 1 lb. flour.
4. 2 oz. fat in 1 lb. flour.
5. 1.8 oz. water in 1 lb. flour.
6. 1 oz. minerals in 1 lb. flour.

EXPERIMENT 19

Purpose.

To show the presence of oxidizable nutrients in foods.

Directions.

a. Flour.

Add to a large spoonful of flour a very little water and make a stiff paste. Place in a square of cheesecloth, and tie the corners together, forming a bag. Place in beaker or glass crystallizing dish of water, and knead the contents of the bag for five minutes. Allow the solid in the beaker to settle, and then pour off the water. Describe this substance.

Continue kneading the bag in a stream of cold water from the faucet until no more comes through the cloth. Open the bag and describe the contents.

NOTE. The substance in the dish is starch, a carbohydrate. The substance in the bag is gluten, found in the kernel of wheat, and belonging to the class of proteins. It is the sticky, elastic nature of this substance which makes it possible to raise dough by means of the gas generated during the action of yeast on sugar, as in Experiment 21, or in the reaction which occurs when baking powder is moistened, as in the experiment found in the supplement. It is probable that much bread is made indigestible by too much kneading, which makes the gluten tough, as in this experiment, and prevents the starch from absorbing water, as it must do to undergo digestion. The use of a bread mixer removes this disadvantage.

b. Butter.

Place a good-sized piece of butter in a test-tube, put the test-tube in water at about 60° C., and allow it to stand until water is evident at the bottom. Describe the appearance of the butter.

NOTE. The yellow substance at the top is a fat; at the bottom is water, holding salt in solution and casein, a protein, in suspension.

c. Sugar in fruit.

Cut into pieces a few dates, place in hot water in a mortar, and grind them. Add considerable water and let them soak a few minutes. Filter the water into an evaporating dish and evaporate over a small flame. Describe what remains. Taste it. To what class of nutrients does it belong?

Test-tube of melted butter showing fat, curd or casein, and water

There are several ways in which carbohydrates, fats and proteins differ from inorganic or mineral substances. In the first place, they are not found in nature except in plants and animals, or as the result of their life processes. They received the name organic substances early in their investigation because of this fact. Later it was discovered that many of them could be made by artificial means, but nature makes them only by means of the living activities of organisms. Typical carbohydrates are the different kinds of sugars found in fruits, beets and the sugar cane, and starches found in potatoes and grains. Typical fats are the yolk of the egg, and oils from the olive, the cocoanut and numerous other plants. Typical proteins are the white of egg, substances in the lean of meat, and gelatine. These illustrations are sufficient to make us realize that such substances present decided differences from foods like table salt and calcium phosphate. Their mode of occurrence in nature is the first distinguishing characteristic.

The second evident difference is the complexity of their composition, and this is especially true of proteins. It will be remembered that in the chapter on carbon compounds the complexity of these substances, and the structural arrangement of atoms in the molecule were emphasized. Not all such carbon compounds serve as foods, but in all oxidizable foods we find carbon, and its high valence, and power to form chains of carbon atoms, which were emphasized earlier, are responsible for much of the complexity of composition found in these foods. The study and even manufacture of these nutrients in the laboratory has proved that in many of them the molecule is very large, and the arrangement of atoms complicated.

A third point of difference between them and other substances is again related to the size of the molecule. So far we have spoken of substances as soluble or insoluble. Sugar and salt are soluble, starch and silver chloride are insoluble. These terms may be used absolutely only for convenience. We can dissolve so much salt or sugar in water that no more will dissolve; we are probably able by raising the temperature or other means to dissolve a small quantity of even the most insoluble substance. More than this, many substances exist in water or other media in a condition between solution and suspension. This condition is called a colloidal condition, and in this state the particles are usually smaller than those suspended and larger than those dissolved. We have already had illustrations of the colloidal state in the hydroxides of iron and aluminum, precipitated in Experiment 7. Now while the colloidal state is not confined to food substances, yet many substances used as foods, particularly among proteins and carbohydrates, occur naturally in this condition, the white of egg for instance, or are put into this state by heating with water, as is the case with starch. Substances in a colloidal state pass through animal membranes like those that line the digestive tract more readily than suspended substances but much less readily than dissolved substances, and so the colloidal condition of foods is of importance in digestion. Indeed the main part which cooking plays in preparing food for eating is in the effect on solubility.

With this suggestion of the differences to be expected in our study of oxidizable nutrients, we will begin with carbohydrates. It will be well to keep in mind, in the experiments that follow, that we are concerned with these substances as foods, and that their condition of solubility and the effect of heat on them, as in cooking, are of primary importance.

QUESTIONS

1. Why, in Experiment 19 *a*, did the starch go through the cloth and the gluten remain behind?
2. Did the butter in Experiment 19 *b* undergo a chemical change? Is it a chemical compound?
3. What would be the characteristics of "colloidal gold"? "colloidal leather"?

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CHAPTER VI

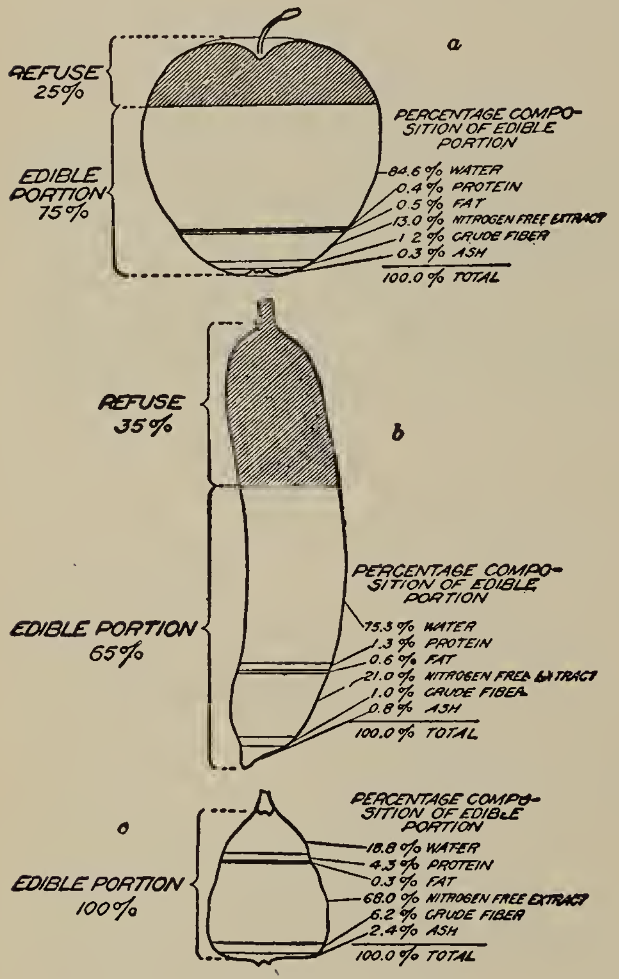
CARBOHYDRATES

Typical Carbohydrates

1. Cellulose as absorbent cotton. 2. Starch. 3. Dextrin. 4. Cane sugar. 5. Grape sugar or glucose.

CARBOHYDRATES include the familiar substances sugar and starch. Intermediate between these two groups comes the class called dextrins or gums, substances formed to some extent in the cooking and digestion of starch, while after starches comes the carbohydrate cellulose, which is found in large quantities in vegetable foods making up their cell walls, but which cannot be digested, and so cannot serve as nourishment to the body. There is a gradual gradation in solubility among the carbohydrates which is interesting as showing their relationships and important in considering any of these substances as foods. Sugars are all very soluble; dextrins and gums tend to exhibit the colloidal condition in water; starches are insoluble in cold water; cellulose is insoluble in boiling water.

There are many sugars, differing in their origin and the size and structure of the molecule, but all having nearly the same percentage composition of elements, and serving the same uses in the body. The most common sugars are classified as monosaccharides, those having the simplest molecular formulas, and disaccharides or those of larger molecules. Ordinary white sugar, obtained commercially from the sugar cane or beet, is called sucrose and is a disaccharide. The sugars found in fruits belong to the monosaccharides, or simpler sugars. All the common sugars, with the exception of sucrose, are reducing in their action. There is a standard test for the presence of these sugars.



From "Use of Fruit as Food," U. S. Farmers' Bulletin 293.

Composition of apple (*a*), banana (*b*), and dried fig (*c*).

EXPERIMENT 20

Purpose.

To study the test for a reducing sugar.

Directions.

Use in equal amounts the two standard solutions which together make up Fehling's solution. They contain copper sulphate and sodium hydroxide and Rochelle salt (sodium potassium tartrate).

Dissolve a little glucose or a sugar obtained from any fruit, as in Experiment 19 *c*, in half a test-tube of water, and bring to the boiling point. Add 1 or 2 c. c. of this solution to 10 c. c. of Fehling's solution. Boil a few minutes. If no red or orange color results, add a little more sugar solution and boil again. Cool. Describe.

This red precipitate is a test for the presence of a reducing sugar.

NOTE. The following reactions take place:

$$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 + Na_2SO_4$$
$$Cu(OH)_2 \rightarrow CuO + H_2O.$$

Cupric oxide forms soluble compounds with the substances present. The sugar on boiling reduces this oxide to the cuprous oxide,

$$2CuO \rightarrow Cu_2O + O;$$

and cuprous oxide, insoluble in Rochelle salt, is the red precipitate obtained. This reducing action is not performed by cane sugar, and so this is not a test for its presence.

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Planting Sugar Cane

Sugars are found in all fruits, as the monosaccharides glucose[1] and fructose; in the sap of plants, especially the sugar cane and the sugar maple, as the disaccharide sucrose; in milk, as the disaccharide lactose; and in many vegetables. We shall study later their relation to starches and shall then see how it is possible for the plant to transform its storage of insoluble starch into soluble sugar, so that it may be carried from root and tubercle to all parts of the growing plant when spring awakens its activities.

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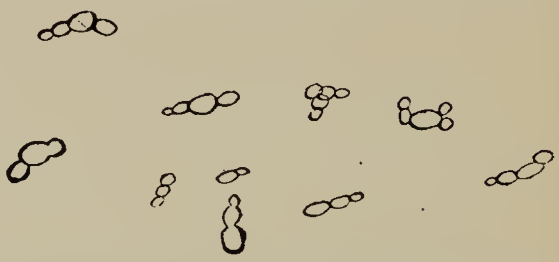
[1] Chemical glucose, or grape sugar, should be carefully distinguished from commercial glucose which is a mixture of dextrose, dextrin, and maltose, and is made by heating cornstarch.

Yeast Cells

The following experiment on the fermentation of sugar will illustrate both the process of making beer and wines from barley and fruits, and also the reason why yeast is used in raising bread. Similar and very undesirable fermentation processes occur in the stomach and intestines if too much sugar is taken in the food.



Yeast cells from the same source after growing two days in a warm molasses solution.

EXPERIMENT 21

Purpose.

A study of fermentation.

Directions.

Mix a piece of yeast cake about the size of a large pea with a few drops of water until a thin paste is made. Mix in a beaker 200 c. c. of warm water and 25 c. c. of molasses, and add the yeast. Place in a flask or test-tube fitted with a stopper through which a delivery tube passes into test-tube full of water. Another tube leads out of the test-tube to drain off the water as the gas formed by fermentation in the flask forces it out. Have all connections very tight. Place a beaker to catch the water. Leave in a warm place 70° – 90° F. over night.

(Two pupils may leave the flasks or beakers open to the air.)

The next day examine contents of flask, state of yeast, and the test-tubes. Test the gas in them with limewater. What is it?

Examine yeast under microscope before and after fermentation. What difference do you see?

Lecture Table.

Distill the contents of flasks at 78° C. (a water bath and thermometer). Re-distill the distillate and test first few drops by placing a match to it. Result? What flame does this resemble?

NOTE. The yeast in growing manufactures a substance called zymase, a ferment. This acts catalytically on the sugar and decomposes it into carbon dioxide and alcohol. This is fermentation.

$$C_6H_{12}O_6 \rightarrow 2CO_2 + 2C_2H_5OH.$$

Starches are the great storage material of plants. Made by combining the carbon dioxide of the air with water from the soil, they are stored in the roots or stem until needed to furnish nourishment for growth. The animal eats this storage material, being unable to manufacture it for himself as the plant does. The following experiment will study starch in its relation to water, and a standard test for its presence.

EXPERIMENT 22

Purpose.

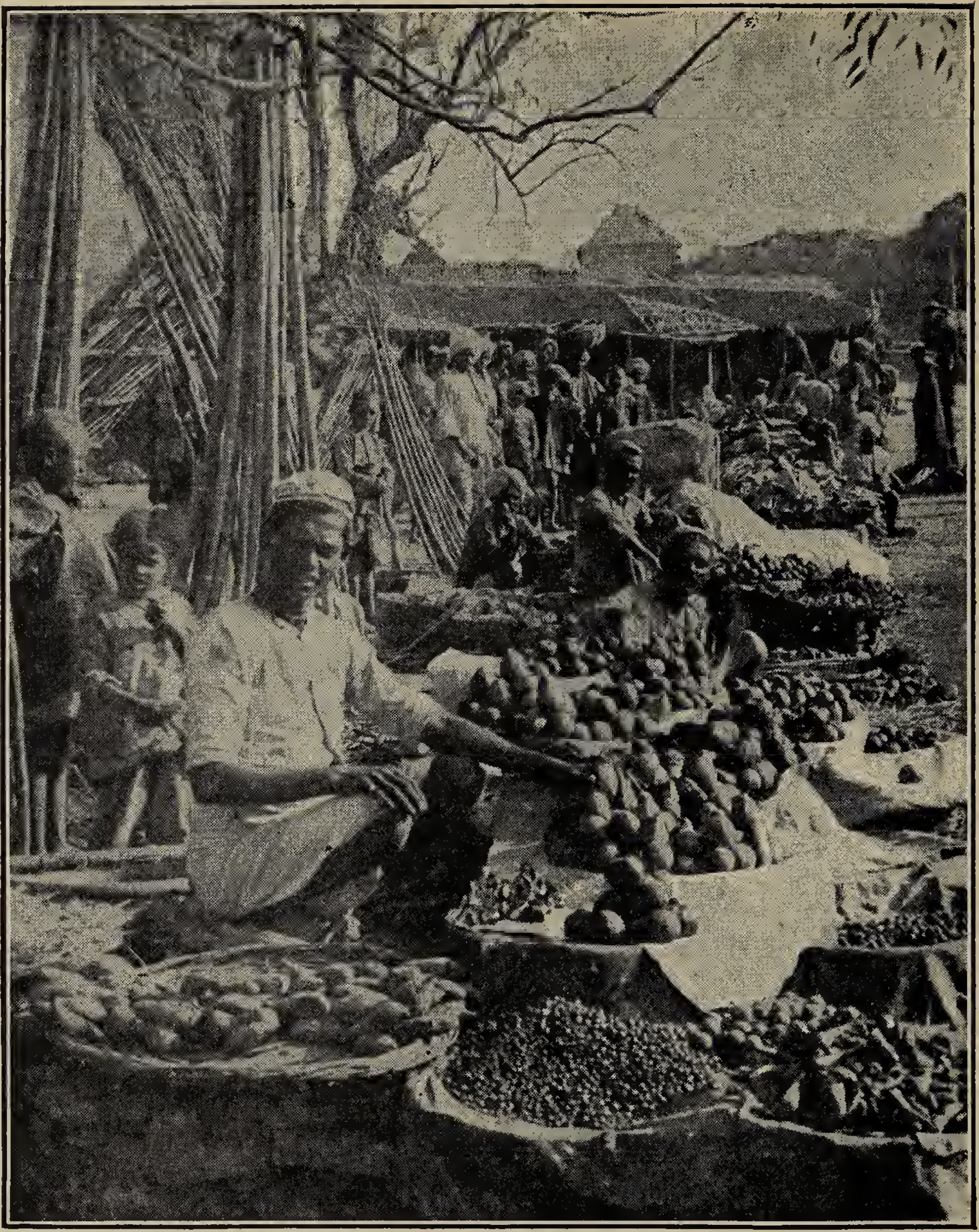
To study the solubility of starch, and a test for its presence.

Directions.

a. Examine a little starch in water under the microscope, noting the grains of which it is composed.

b. Mix a little powdered starch with a little water, making a paste. Add this to half a test-tube of water, and shake. Does it dissolve? Heat gently. Describe what happens and the appearance of the substance. Heat the slide containing the starch grains and again examine under the microscope. What has become of the grains?

c. Add a very little of the starch paste made in (*b*) to half a test-tube of water, and add a drop of tincture of



Ⓒ Underwood & Underwood, N. Y.

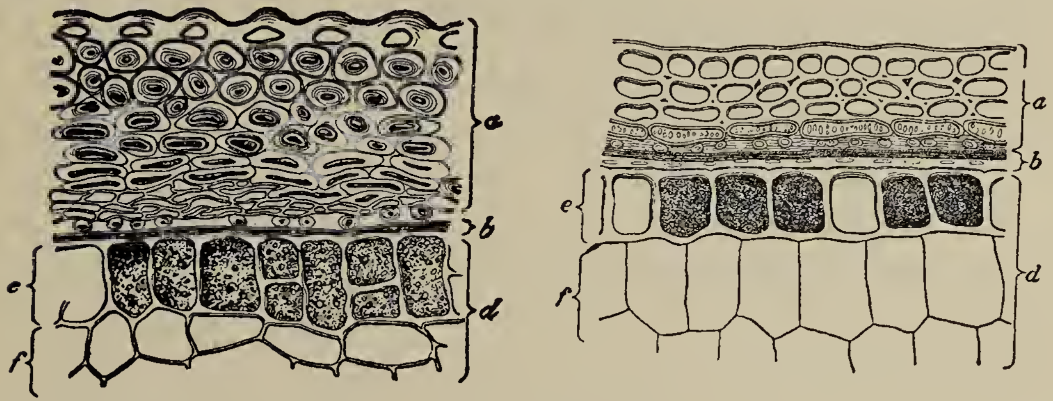
Vegetables for Sale in India

iodine. Result? This blue color is a test for the presence of starch. Heat gently. Result? Is iodine a test for hot starch? Allow to cool. Result? Now boil hard for a few minutes, and again cool. Result?

d. Test potato, rice and white of egg for starch. Which contain it and which do not? Is water necessary to this test? Is it a test for uncooked or cooked starch, or both?

Questions. Is starch soluble? What is its condition as to solubility after heating in water? Is cooking, then, desirable for starchy foods? Does starch need to be digested after being cooked? Give reason for your answers.

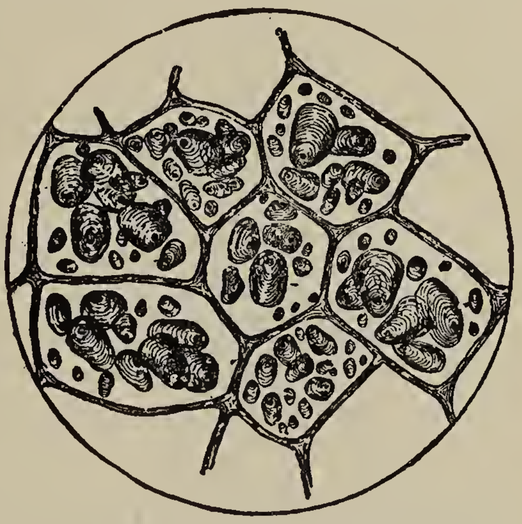
The foods containing most sugar have already been mentioned. Starchy foods are chiefly the grains, vegetables and fruits. The tables of food values should be consulted at this point, and the percent of carbohydrates in the common foods noted. Among the vegetables potatoes and the ripe legumes, peas and beans, are particularly rich in carbohydrates, while among the cereal products rice, hominy and rye flour stand high, although all cereals contain large percents of starch.



From "Bread and Bread Making," U. S. Farmers' Bulletin 389

Cellular Structure of Corn and Wheat
(after Winton and Moeller)

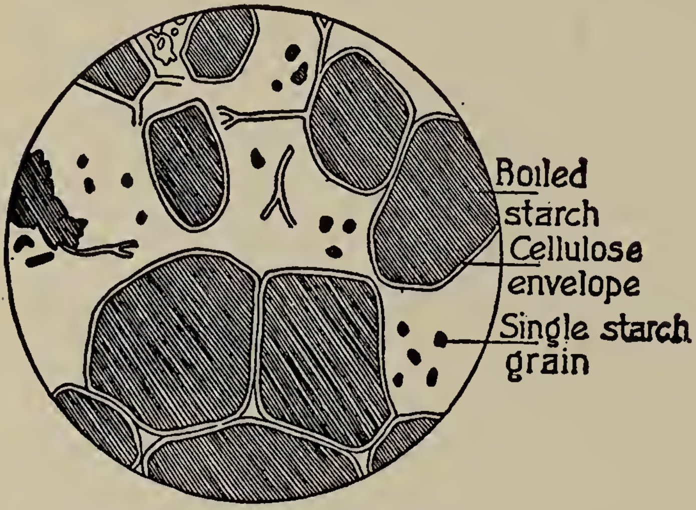
(*a*) Skin and testa; (*b*) membrane; (*d*) endosperm, consisting of aleurone cells (*e*) containing protein, and starch cells (*f*).



From "Principles of Physiology and Hygiene" By G. W. Fitz, M.D.; published by Henry Holt and Company

Cells of a raw potato, with starch grains in natural condition.

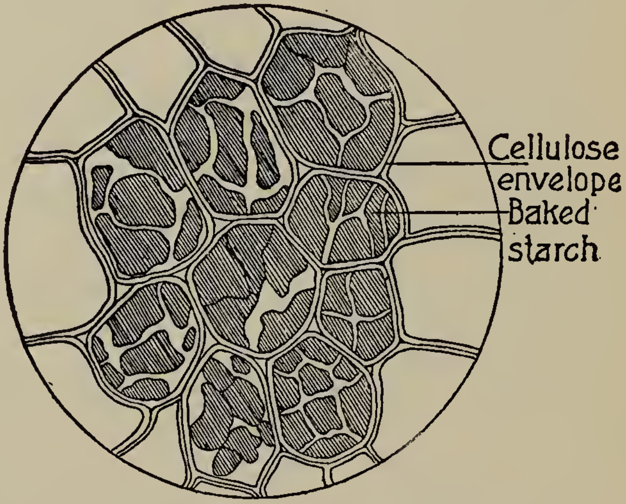
The previous experiment has illustrated the need for thorough cooking of all these foods. If vegetables and cereals are not thoroughly cooked by a slow, moderate heat so that they are well softened, the starch in them remains insoluble, and is lost to the body as food. Besides this, while cooking does not dissolve the cellulose that makes up the walls of the cells in which the starch grains are found, it does soften it so that it is more easily removed by the physical processes of digestion, and so that the starch is more readily exposed to the action of the digestive juices.



From "Principles of Physiology and Hygiene"
By G. W. Fitz, M.D.; published by Henry Holt and Company

Mashed boiled potato. (Bulletin of Agricultural Department.)

The dextrins are not sufficiently important as food to deserve a long discussion. Some dextrin is formed from starch in flour during the rising of bread, by the action of an enzyme in the yeast. It is also formed by heat in the crust of bread, and when bread is toasted. It is of course more nearly ready for absorption than starch, being half way between starch and sugar in solubility, and for this reason dried and toasted bread is more digestible than freshly baked bread.



From "Principles of Physiology and Hygiene"
By G. W. Fitz, M.D.; published by Henry Holt and Company

Baked potato, showing masses of cooked starch within envelopes of cellulose. (Bulletin of Agricultural Department.)

Another related substance is caramel. This is really a mixture of compounds. It is formed when sugar is heated, by decomposition and loss of water. Caramel has a dark brown color and a bitter taste, and is often formed during cooking operations.

EXPERIMENT 23

Purpose.

To study the relationship among carbohydrates.

Directions.

Materials: Glucose (grape sugar), cane sugar, dextrin (use pure, white variety), starch, cellulose in the form of absorbent cotton or filter paper.

Taste each substance and name those that have a sweet taste. Are the sugars equally sweet?

Weigh half a gram of each and test its solubility by stirring it with 15 c. c. of water in a small beaker. Name the soluble ones. Which sugar is less soluble?

Boil the contents of each beaker. In which is a change evident? What is it?

Test a few c. c. from each with Fehling's solution for a reducing sugar. Which gives the reaction? Dextrin may give the reaction if impure.

To the solution of cane sugar (sucrose, a disaccharide) add a little concentrated hydrochloric acid and boil for a minute or two. Cool. Add a little litmus solution, and then sodium hydroxide solution drop by drop until the acid is neutralized. Filter if necessary, add a little Fehling's solution, and boil. Result?

Repeat with a solution of dextrin in water, using concentrated sulphuric acid, and boiling about twice as long. Neutralize with solid sodium carbonate, filter if necessary, and test with Fehling's solution. Result?

Repeat with the starch, boiling with acid until a few drops of the cooled solution added to iodine solution no longer produce a reaction for starch. Boil a little longer to be sure that the reaction is complete. Continue as with dextrin.

Pour off the water from the cellulose, cover with concentrated sulphuric acid, and allow to stand a minute or two. Does the cellulose dissolve? Add very carefully five or six times its volume of cold water, and boil for two or three minutes. Neutralize with sodium carbonate and test with Fehling's solution. Result?

Question. Into what form may all the carbohydrates be changed by boiling with dilute acid?

This experiment makes plain the fact that these substances are closely related in chemical composition. The acid serves as a catalytic agent while the action is really one between the carbohydrate and the water — an action which is called hydrolysis. The relationship between the various sub-classes of carbohydrates is thus established. We know, too, that only in their soluble form can substances be absorbed into the blood and so help to nourish the body. The question then remains, to what extent may such a process take place in the body? The following experiment will answer this question.

EXPERIMENT 24[1]

Purpose.

To study the digestion of starch.

Directions.

Make a little starch paste as in Experiment 22. Cool to about 40° C. Add two or three drops of sodium bicarbonate solution, and stir. Place a few c. c. in each of two test-tubes, labeled *a* and *b*. Collect 2 or 3 c. c. of saliva for *a*, and mix a pinch of pancreatin or of amylopsin with 5 c. c. of water for *b*.

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[1] The digestive enzymes can be purchased as powders from any wholesale druggist, and keep satisfactorily in this form if tightly stoppered. The two essential ones for the following experiments are pepsin and pancreatin. These can both be obtained in one ounce bottles from Fairchild Bros. and Foster, New York. They also supply trypsin in one-fourth ounce bottles. Amylopsin in one ounce bottles can be obtained from John Wyeth and Bro., Philadelphia.

Have ready several test-tubes, each containing 5 c. c. of a solution of tincture of iodine — two drops of the tincture in 60 c. c. of water makes a good testing solution. Have ready also several test-tubes containing Fehling's solution, 5 c. c. of each part in a tube. Have ready a water bath at about 40° C. — water heated in a small beaker does very well. The temperature should be between 35° and 40° C., not above 40° C.

Before adding the saliva or pancreatin test the starch paste by putting a drop into the iodine solution, and by boiling 1 or 2 c. c. with Fehling's solution. Then add the saliva to *a* and the pancreatin or amylopsin to *b*, shake thoroughly, and place the tubes in the bath. At the end of five minutes test again for starch and for sugar, and again in five minutes if necessary. Results? Compare the results with the first tests. Is the starch as strong? Is sugar present? What difference is there in the appearance of tubes *a* and *b* now and before the ferments were added? What has happened to the starch?

NOTE. The sugar formed in this instance is maltose, a disaccharide.

The starches are acted upon in the mouth, and this action is continued for some time in the stomach by ptyalin, a substance found in the saliva. Ptyalin is the first example of enzymes, a class of substances found in saliva, in gastric juice in the stomach, in pancreatic and intestinal juices in the small intestines. These enzymes, or unorganized ferments as they are sometimes called, are chemical substances whose composition is unknown. They act by catalysis, that is, they cause reactions without themselves taking part, just as manganese dioxide caused the potassium chlorate to decompose in the experiment for obtaining oxygen. The general type of reaction which takes place in digestion consists in a combination of the nutrient with water, and then a splitting of this molecule into two smaller ones of a new and simpler substance. This may be repeated many times, until finally the molecules are ready for absorption into lymph or blood. The action is one between the molecules of the nutrient and those of water, and so is called hydrolysis. The enzyme causes the action, different enzymes being necessary for different nutrients. In the case of starches, the same action is caused by both the ptyalin of the saliva and the amylopsin of the pancreatic juice. This latter juice is poured into the small intestines by a duct which brings it from the pancreas, a large gland that secretes or manufactures it from the blood. Both ptyalin and amylopsin act in an alkaline solution, the food being alkaline in the mouth, acid in the stomach, and again alkaline in the intestines.

The digestion of carbohydrates is mainly the digestion of starch. Disaccharides such as cane sugar, however, although soluble, must also be hydrolyzed. They are split up into monosaccharides of smaller molecules before they are absorbed. The maltose formed by the action of pytalin and amylopsin on starch, and the sucrose and lactose of other foods are hydrolyzed to some extent in the stomach, but more extensively by the action of the intestinal juice in the small intestines. Very little tax is placed on digestion by sugars, and they are soon received into the capillaries of the stomach and intestines. As a means of quick restoration from an exhausted condition sugar works rapidly and is valuable. Its excessive use in place of other nutrients, particularly if eaten, as candy so often is, in large quantities and without other food, is not good. As has been already stated, sugars are particularly subject during digestion to fermentation caused by bacteria, which breaks them up into acid substances not useful as food, and causes irritation of the digestive tract.

We have still to consider the chemical composition of carbohydrates and the use which the body makes of them. Enough has been shown in the experiments to demonstrate the close relationships among all these substances. They all contain only the elements carbon, hydrogen and oxygen, and always the two latter elements in the same proportion, that found in water, two atoms of hydrogen to every one of oxygen. The simplest sugars, the monosaccharides, have the formula $C_6H_{12}O_6$, differing in structure but alike in percentage composition. The disaccharides have the formula $C_{12}H_{22}O_{11}$. The simplest formula for the dextrins, starches, and cellulose is $C_6H_{10}O_5$, any particular number having some multiple of this for its molecular weight. Thus $(C_6H_{10}O_5)_n$ may be taken as their general formula. Many of these substances have thousands of atoms in the molecule, and their complicated structure may be imagined. The complete hydrolysis of starch may now be represented in symbols.

$$(C_6H_{10}O_5)_n + nH_2O \rightarrow nC_6H_{12}O_6.$$

That for cane sugar is

$$C_{12}H_{22}O_{11} + H_2O \rightarrow 2C_6H_{12}O_6.$$

The manufacture of starch by plants from carbon dioxide and water is symbolized thus:

$$6CO_2 + 5H_2O \rightarrow C_6H_{10}O_5 + 6O_2.$$

A very entertaining as well as instructive demonstration of the proportions of hydrogen and oxygen in the chemical composition of the carbohydrates may be made by the action of sulphuric acid on sugar. This acid combines greedily with water, many of its uses depending on this dehydrating action.[1] In a large beaker, about 500 c. c., mix thoroughly 60 grams of lump sugar with 45 c. c. of warm water, making a thick syrup. Place the beaker on a plate, and pour in 60 c. c. of concentrated sulphuric acid. The result is seen in the illustration.

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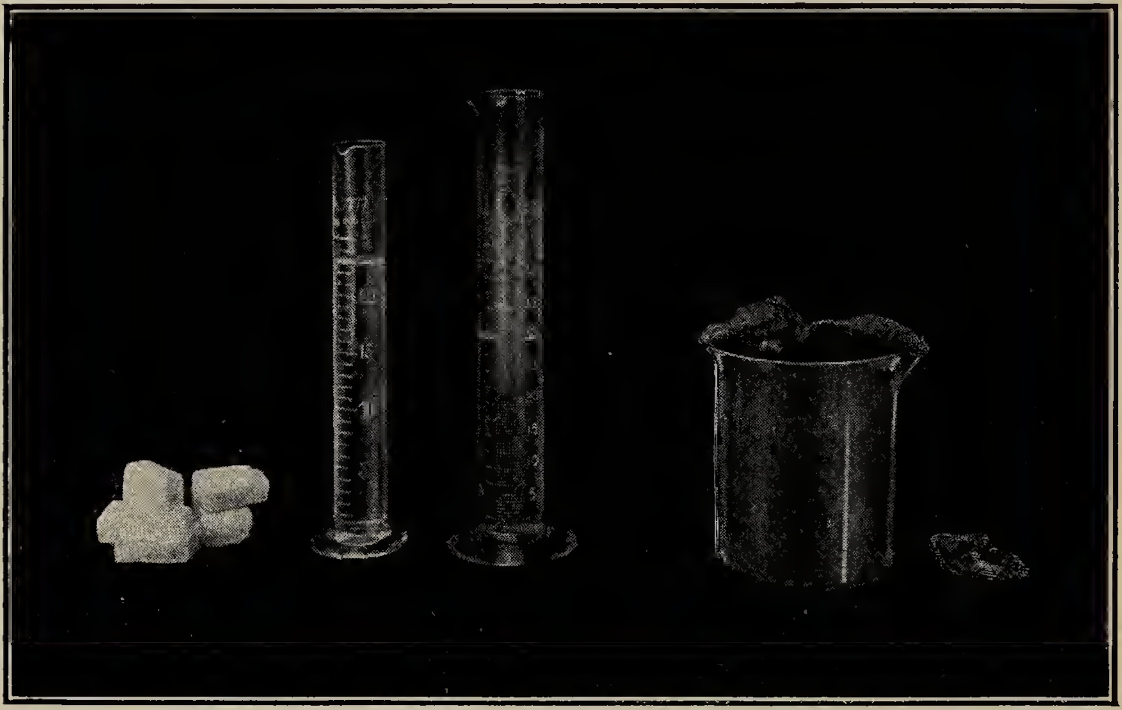
[1] See F. G. Benedict's "Chemical Lecture Experiments," p. 175.

The use made of carbohydrates in the body is twofold. They are not used to build up living tissue, but they do furnish heat and energy by their oxidation, as do the proteins, and to the same extent, one gram of each of these nutrients furnishing four Calories of heat. They have a further use in that they are manufactured into storage material which the body may call on later in case of need. This storage material is found in two forms.



The carbonization of sugar by sulphuric acid. 30 g. of lump sugar, 22.5 c. c. of warm water, 30 c. c. of concentrated acid.

A carbohydrate named glycogen, often called animal starch, is stored in the liver and to a less extent in the skeletal muscles. When conditions in the body demand it, this glycogen is transformed into sugar, and is carried by the blood in this soluble form to any part of the body where it may be needed, just as the plant transforms its starch into sugar and carries it dissolved in sap up to leaves and flowers. On the death of an animal, glycogen undergoes this change into sugar, which explains the sweetish taste of the liver used as food. Carbohydrate food is the chief source of supply for glycogen, but it may also be formed from protein, and probably from fat.

The second form in which material may be stored in the body is as fat. The chief places for such storage are the cells of connective tissue, especially that directly under the skin. Here the protoplasm of the cells is often pushed aside to make room for droplets of oil which almost fill the cell. Tissue which thus stores fat is called adipose tissue. The carobhydrates in the food furnish a large part of the material transformed into the fatty tissue of the body.

Cellulose in vegetables is almost entirely useless as a food, since the digestive juices cannot render it soluble. It is very important in digestion, however, because the coarseness of its texture stimulates by mechanical action the lining of the intestines, and so the muscular movements of that organ are quickened, digestion is facilitated, and the danger of constipation is lessened. This is one of the chief reasons why bulky vegetable foods, often of small calorific value, are still indispensable in the diet.

QUESTIONS

1. Judging from Experiment 20, in what class of substances, grouped according to their chemical action on other substances, may sugar be placed? Do you think that its action in this respect is vigorous or not? See the directions for making a mirror referred to in the Supplement.
2. Why is the sap of trees sweetish in taste? What connection is there between this fact and the fact that gums are made from some saps?
3. Why is vinegar used in the making of candy to be pulled?

4. Does adding sugar to lemonade neutralize the acid in it?
5. Why is baking soda added to tomatoes while stewing? Why is sugar added?
6. Why is dry and toasted bread more digestible than fresh bread?
7. Why is ptyalin called an unorganized ferment while yeast is not?
8. Why did the cellulose darken when sulphuric acid was poured on it in Experiment 23?

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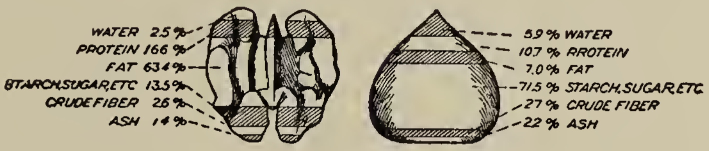
CHAPTER VII

FATS

FATS occur in many foods, and are an important class of nutrients. Meat, eggs, milk, and its products butter and cheese, nuts and other seeds, such as flaxseed and cottonseed, are the chief natural products containing large proportions of fats and oils. Vegetables and fruits contain very little fat, if any.



From "Nuts and Their Uses as Food," U. S. Farmers' Bulletin 332

Percentage composition of an oily nut (walnut) and a starchy nut (chestnut).

EXPERIMENT 25

Purpose.

To study a test for fats and oils.

Directions.

Use yolk of egg, olive oil, suet or any meat fat, flour, glucose, starch, nuts — preferably walnuts. Label a test-tube or small dish for each food, cut each into small pieces or grind in a mortar if necessary, and place a small quantity in the tube. Pour on enough ether to more than cover the substance. Shake or stir, and allow to stand a few minutes.

Rub a glass plate thoroughly clean, and pour on it a few drops from the clear liquid in each tube. Filtering is not desirable. Compare the marks left on the glass by the evaporation of the ether. Ether is a solvent for fats and oils, and on evaporation these substances remain on the glass in a greasy film. If in doubt as to whether the residue is greasy or not, rub the spot gently with the finger tip, and any greasy nature will be at once evident.

Questions. Which of these foods contain fat or oil? Which do not?

The fats and oils as they occur naturally in foods are usually mixtures of different fats. This may be shown by placing some cottonseed oil or olive oil in a test-tube and surrounding it by a mixture of ice and salt. When thoroughly cooled, a small quantity of white solid will separate out. This is a fat of definite composition called stearin. Thus olive oil is shown to be a mixture of fats and not one chemical compound. If the oil be warmed the stearin again goes into solution in the oil, but the composition is evidently not just the same as before, for the oil will not now make so good a salad dressing. Chemically pure fats are colorless, tasteless and odorless. They are not found so in foods, for small quantities of other substances give them color, taste and odor.

The chemical composition of fats is similar to that of carbohydrates, although fats are more complex in structure. They contain only the elements carbon, hydrogen and oxygen, but in different proportions from those of carbohydrates. In fats there is much less oxygen in proportion to the carbon and hydrogen. Fats and oils are all esters, or organic salts of the organic base glycerine, or glycerol, and various fatty acids. The most common fatty acids forming fats used as foods are stearic acid, $H(C_{18}H_{35}O_2)$, a soft solid; palmitic acid, $H(C_{16}H_{31}O_2)$, also a soft solid; and oleic acid, $H(C_{18}H_{33}O_2)$, a liquid. The formula for the organic base glycerine is $C_3H_5(OH)_3$. The esters formed by the union of these substances would be glycerol stearate, often called stearin, $C_3H_5(C_{18}H_{35}O_2)_3$, seen when olive oil was cooled; glycerol palmitate, or palmitin, $C_3H_5(C_{16}H_{31}O_2)_3$; and glycerol oleate, or olein, $C_3H_5(C_{18}H_{33}O_2)_3$. These substances are fats, the first two solids and the last an oil. As we have seen in the case of olive oil, these usually occur mixed together or in loose chemical combination, and the physical condition of any fat as a firm solid, a soft solid, or a liquid, depends on the different proportions of these ingredients. Suet contains more stearin and palmitin than does lard, and lard more than do the oils.

The glycerol and the fatty acid which make up the fat are in loose chemical combination and are easily separated when there is present a substance that is capable of bringing about this change. Here again we have the hydrolytic action already familiar in the experiments with carbohydrates. The esters combine with water and then split up into the base and acid which originally formed them. Bacterial action, or the presence of certain enzymes, is sufficient to produce this change. When fats become rancid it is because certain fatty acids are formed which have this rancid nature.

EXPERIMENT 26

Purpose.

To test for fatty acids.

Directions.

Add a drop of phenolphthalein to about 10 c. c. of alcohol in a test-tube. Add a diluted solution of sodium hydroxide drop by drop until the alcohol is colored red. Phenolphthalein is an indicator like litmus, colorless in neutral or acid solutions but red in alkaline solution. Introduce a few cubic centimeters of fresh olive oil, butter, lard, or other fat, and shake vigorously. Repeat with a second mixture, using a fat suspected of being rancid. Allow both test-tubes to stand for some time. If the fat or oil is at all rancid, the red color will gradually fade. Why?

NOTE. All fats probably contain small amounts of free fatty acid, and are not therefore necessarily rancid; but if a fat is rancid, it is always acid as well.

A recent industrial development is worth mentioning here because by means of it large quantities of fats that would otherwise be useless as food are made palatable. This is called hydrogenation, and consists in forcing hydrogen through cottonseed oil in which is suspended a catalytic agent such as finely divided nickel. The oil, being what is called unsaturated, enters into combination with the hydrogen, forming a saturated fat which is a solid, and which is then sold as a lard substitute. In a similar way fish oil loses its offensive odor and becomes useful in making soap.

EXPERIMENT 27

Purpose.

To show the results of mixing oils with water and with water containing certain other substances.

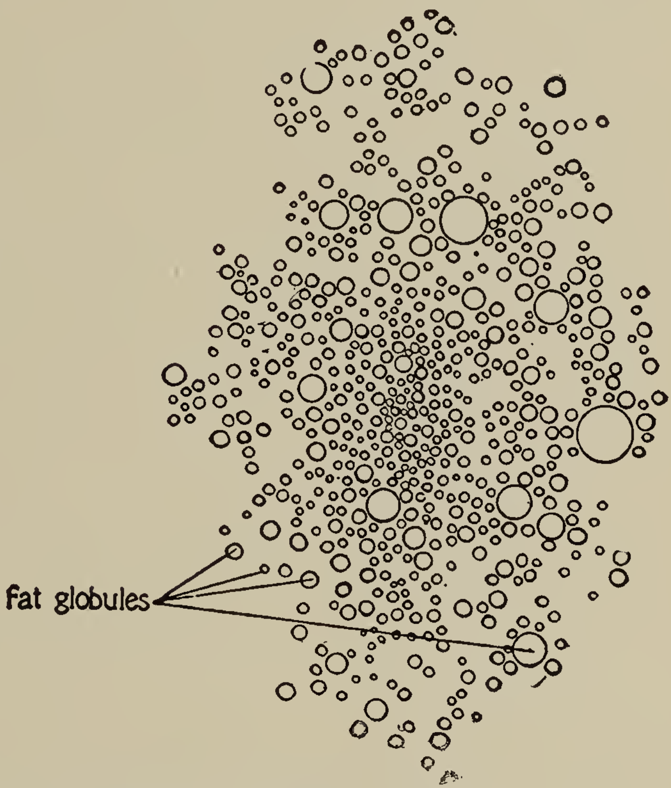
Directions.

Place in each of four test-tubes some cottonseed or olive oil to the depth of one-fourth inch. Label (a), (b), (c) and (d).

To (a) add water to the depth of an inch, to (b) soap solution, to (c) water containing a drop of Na_2CO_3 solution, to (d) albumin solution, made by beating together equal volumes of water and uncooked white of egg.

Shake each test-tube and allow to stand. What difference is evident between (a) and the other three? In (b), (c) and (d) an emulsion has been formed. Which makes the best emulsion? What common food does this look like?

An emulsion is a permanent mixture of an oil with a liquid, the oil being suspended in very tiny drops throughout the liquid, and not settling out on standing. It is evident that water will not make an emulsion with fats and oils, but that if certain substances such as soap or soluble protein are present, an emulsion may be formed. The explanation is that the soap or the albumin collects around each tiny droplet into which the oil is broken by the shaking, and this thin film of substance, in a colloidal state, prevents the droplets of oil from combining and so makes the finely divided condition of the oil permanent. Milk is such an emulsion, as may be seen by putting a drop under the microscope, when the tiny particles of fat are evident. A protein in the milk called casein performs the work of the soap or albumin in the experiment. The rising of cream and the making of butter both show the separation of the lighter fat particles from the water, sugar and salts of the milk, some of the casein remaining in the liquid and some separating out with the fat. This property of forming emulsions is important in digestion.



From "Principles of Physiology and Hygiene" By G.W. Fitz, M.D.; published by Henry Holt and Company

A portion of a drop of cow's milk highly magnified, showing the fat globules.

It is not practicable to study the digestion of fats in an elementary laboratory. This work is carried on in the small intestine by the pancreatic juice, aided by the bile. The pancreatic juice is the most important digestive secretion. The action of its enzyme amylopsin on starches has already been studied in Experiment 24. Another enzyme has the power to digest protein, as we shall see in the next chapter. This juice has also the power of emulsifying fats because it contains protein substances which surround the droplets as did the white of egg in Experiment 27 *d*. It is possible that under certain conditions these microscopic droplets may pass unchanged through the lining of the digestive tract, perhaps squeezing in between the cells. As a rule, however, fats are hydrolyzed before being absorbed. The pancreatic juice contains an enzyme for this work also, called a lipase, or fat-splitting ferment. This steapsin, as it is sometimes named, acts catalytically, as do all enzymes, to cause the hydrolysis of fats, which combine with water and then split up into glycerine and a fatty acid. A similar lipase in the gastric juice of the stomach also splits fats, but only when they are taken in an emulsified condition.

$$C_3H_5(C_{18}H_{33}O_2)_3 + 3H_2O \rightarrow C_3H_5(OH)_3 + 3H(C_{18}H_{35}O_2)$$

These substances, glycerine and fatty acids, are now ready for absorption. They are taken into the lymphatics in the lining membrane of the small intestine, and thence pass into the blood. During absorption they are reconverted into fats, and appear in this condition in the blood.

Bile is a juice manufactured by the liver, stored in the bile-sac, and poured into the small intestine with the pancreatic juice when digestion begins. It is important because it aids the digestion of all three classes of nutrients. It seems to emulsify fats, and in it fatty acids, which are not soluble in water, are dissolved and so may be absorbed. It also contains waste products which are to be excreted. It is important in regulating the action of the intestines and preventing constipation.

A juice called the intestinal juice is poured into the small intestine throughout its course from tiny glands in the lining membrane. In addition to enzymes that hydrolyze disaccharide sugars to monosaccharides, and an enzyme that digests protein, it contains in solution sodium carbonate, and so performs the very important function of keeping the food mass alkaline so that the enzymes may do their work. A direct reaction between the alkali and some of the fat present is important and will be understood better when the chemistry of soaps is studied. A small quantity of fat may be saponified or made into soap in the intestines by a reaction like the following:

$$3NaOH + C_3H_5(C_{18}H_{35}O_2)_3 \rightarrow 3Na(C_{18}H_{35}O_2) + C_3H_5(OH)_3$$

This sodium salt of oleic acid is a soap, and would help to emulsify a larger quantity of fat, as the soap solution in Experiment 27 *b* emulsified the olive oil.

EXPERIMENT 28

Purpose.

To study the effect of heat on fats.

Directions.

Work under the hood.

Place a piece of suet or any meat fat in a test-tube and heat gently. What is the first change produced? Increase the heat and boil until the fat darkens. Describe what happens, the appearance and odor of the substance.

Questions. Do you think it desirable to eat fat cooked at high temperature? What general type of reaction seems to be caused by the heat?

NOTE. The crackling noise is due to the fact that steam escapes, condenses, and runs back into the fat. The dense fumes and the odor are due to acrolein, a substance formed by the decomposition of glycerine.

It is evident that a slight degree of warmth, such as that of the digestive tract, melts solid fats and so prepares them for emulsion. It is also evident that fats are readily decomposed by heat. The application of this fact to the cooking of fatty foods is important. Fats must be classed with proteins as nutrients that are likely to be made less digestible, and in the case of fats, even harmful, by cooking. Their decomposition produces substances that are not only unpleasant but injurious in the digestive tract. For each kind of fat there is a point of temperature above which decomposition takes place, and substances are formed that are responsible for the stinging sensations produced on the eyes and nose during the frying of food, and that are also harmful in a like manner to the lining of the digestive tract. Care should be taken to introduce as little of such substances as possible into the body.

When foods such as doughnuts and oysters are fried in fat the fat should be very hot, boiling in fact, when the food is introduced, so that a protective covering is immediately formed around the doughnut or oyster and the fat has no chance to soak into the dough. Certain foods, such as gravies made with flour, and pastry, are especially difficult to digest because they are such intimate mixtures of starch and fat. It has been seen that starch granules must get into contact with water so that they may swell and become soluble. The fat in gravies and pastry so interferes with this action that the saliva has no chance to affect the starch. Fats, we have seen, are not digested until the food reaches the intestines, and it is plain that by interfering so long with the solution of the starch, they throw an added burden on the intestine, and indigestion as well as loss of food material to the body may result. It is well for all people of delicate digestion to avoid such foods. Fried food in general is likely to be less digestible than that cooked by other methods.

Fats are very similar to carbohydrates in their chemical composition, and serve the same uses, except that they are not used in the manufacture of glycogen. The fatty acids and glycerine into which fats are split during digestion are combined again, and appear in the blood and tissues only as fat. The manufacture and storage of fat from starch has already been mentioned, and fats in the food are used also for this purpose by the body. Like carbohydrates they do not build up the active, living tissues. Like both proteins and carbohydrates their oxidation liberates energy for heat and mechanical work. Because they contain so little oxygen they liberate much more energy than do either of the other two classes of oxidizable nutrients, one gram of fat giving a little more than nine Calories of heat, or between two and three times as much as proteins and carbohydrates.

QUESTIONS

1. In what forms do plants store food? In what form do animals store it?
2. What advantage is there in giving an invalid fatty food in the form of emulsions?
3. Why can milk be so easily digested by infants and invalids?
4. Can you tell why churning makes the butter form? Think of the fat as in a colloidal condition in the milk, each particle being surrounded by an envelop of protein.

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CHAPTER VIII

PROTEINS

EXPERIMENT 29

Purpose.

To study a typical protein.

Directions.

a. Solubility in cold and boiling water.

Place about 5 c. c. of uncooked white of egg in a test-tube and add 5 or 10 c. c. of water. Shake thoroughly. Does the white of egg dissolve? Try filtering it. What seems to be its condition?

Heat slowly to boiling. Result? What is the effect of heat on solubility here?

b. Test for protein.

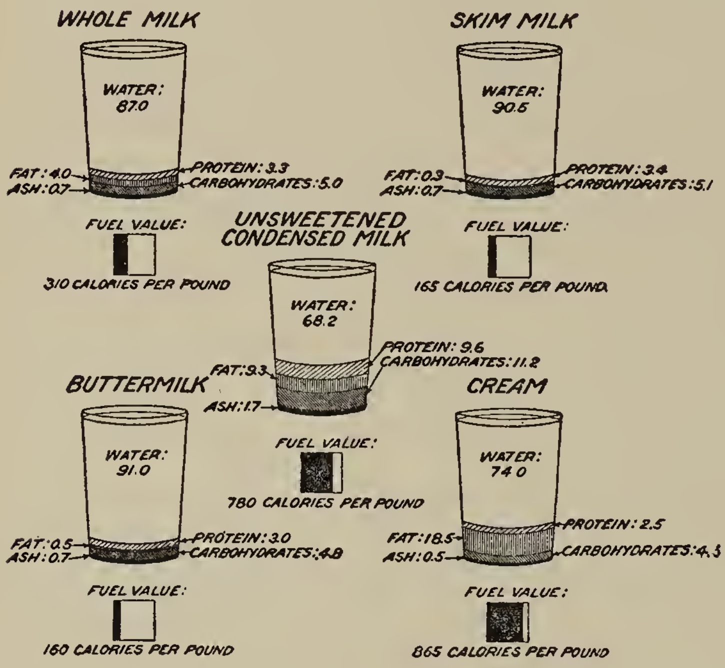
Pour off the water in *a*, and add to the solid a few drops of concentrated nitric acid. Result? Wash off excess of acid with water, and add concentrated ammonium hydroxide. Result? This reaction is typical of proteins in general. Apply this test to a piece of lean meat, beans, peas and potato. Which contain much protein and which do not?

Cooked and uncooked white of egg in water; showing the effect of heat on a typical protein.

It is evident from this experiment that heat decreases the solubility and therefore the digestibility of proteins. The application of this fact to the preparation of foods is obvious. The foods containing most protein come from animals. They are the lean of meat, which contains the proteins fibrin, musculin and others; milk and cheese, containing the casein already demonstrated; and eggs containing albumin. Among vegetable foods wheat flour contains gluten, and peas and beans, or legumes, contain legumin. In the cooking of all these proteins the same principle holds; too great a heat decreases to a decided degree their solubility, and so makes them less easily digested. If proteins existed separately as foods, the problem of cooking them would be easy. It is evident that boiled milk, cooked cheese, hard-boiled eggs, over-done meat, are less easily digested than the uncooked forms. But in many protein-containing or nitrogenous foods the presence of other nutrients makes the problem more difficult.

In bones for instance, and in the connective tissue which forms a network throughout meat, there is present a substance called collagen, belonging to a class of nitrogenous substances closely related to proteins, but affected differently by heat. Heat makes this substance swell in size, and renders it more soluble. Dilute acids have a similar effect on collagen, and this fact explains why tough meat is made more tender by soaking in vinegar. The meat of freshly killed animals is tough, and becomes more tender after a few days because by bacterial action acids are formed in it which soften the collagen. The cooking of meat must be carried on with reference to both collagen and the proteins. If it is to be broiled or roasted, it is exposed to a high temperature at first. This coagulates the protein on the surface, and prevents the juices within from escaping and being lost. Then the temperature should be lowered, during the rest of the time, so that the inside of the meat is cooked at a moderate heat, and the protein within not coagulated. If the meat contains much tough connective tissue, it is well to soften this beforehand, by pounding or by soaking in vinegar.

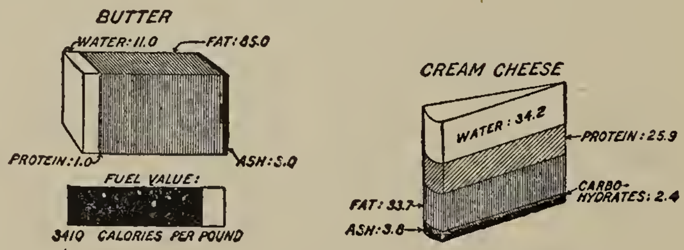
In making stews, the nourishment is desired in the water in which the meat is boiled. The meat should therefore be cut up and soaked in this water first. Then the water should be gradually warmed. The proteins in the meat, more soluble at lower temperatures than that of boiling water, are thus dissolved out. The scum which rises to the top as the simmering goes on consists of this coagulated protein, and forms the most nourishing part of the broth. If the water be finally boiled for a few minutes the collagen is dissolved, and when it cools, shows itself as gelatin, or stock. By this method all the nitrogenous substances may be obtained from meat and bones, and a most nourishing broth results.



From "The Use of Milk as Food," U. S. Farmers' Bulletin 363

Percentage composition and fuel value per pound of whole milk and some milk products.

Whenever, in heating a liquid such as milk, cocoa, stews and soups, a thin scum or skin forms on the top, it is the protein coagulated or made solid by the heat. Evidently such material should not be removed, because this removes much valuable nourishment. It should be stirred or beaten into the liquid.



From "The Use of Milk as Food," U. S. Farmers' Bulletin 363

Percentage composition and fuel value per pound of butter and cream cheese.

In vegetable foods containing proteins, such as wheat flour and peas and beans, there is found a large proportion of starch. As we have seen, this, unlike proteins, is made more digestible by cooking, and so the cooking of such foods is complicated. It may be stated here, however, that because of their proteins vegetables should not be boiled hard, and that they should not be cooked in hard water, or water containing calcium and magnesium salts, because the legumin in them unites with these salts to form insoluble compounds.

The cooking of eggs deserves a word. If they are to be cooked in the shell, they should be placed in water just below its boiling point, covered, and left there as the water cools for five or ten minutes. This makes both yolk and white mealy and prevents the extreme hardening or coagulation which is so undesirable. The old fashioned steamer is perhaps the best method for producing a delicious " soft-boiled " egg.

EXPERIMENT 30

Purpose.

To study the digestion of protein.

Directions.

Press through a sieve the white of a hard-boiled egg. Place one-third of it — about 10 gms. — in a bottle of about 50 c. c. capacity. Add 35 c. c. of distilled water and 1 c. c. of dilute HCl. Rub the egg as fine as possible. Make a solution of pepsin by dissolving .02 gms. in 29 c. c. of distilled water to which has been added 1 c. c. of dilute HCl. Mix thoroughly, and add 5 c. c. of this solution to the white of egg in the bottle. Make a similar test with trypsin or pancreatin, 10 gms. of white of egg, 40 c. c. of distilled water, .02 gms. of trypsin or .06 gms. of pancreatin. Make sure that this reaction is alkaline by adding a few drops of sodium bi-carbonate solution.

Cork the bottles tightly and place in water kept at about

52° C. (126° F.), not higher. It will be found best to heat all the bottles of the students in a large shallow pan at the lecture table, where the temperature may be kept uniform more easily. Shake occasionally and observe the contents. Allow the reaction to proceed for two or three hours. What evidence is there that the albumen has been largely changed into soluble substances? These substances are called proteoses and peptones. Filter the contents of each bottle, and test a portion of the filtrate with zinc sulphate, and another portion by boiling. Repeat these tests with uncooked white of egg in water to see the difference.

The gastric juice contains a very little hydrochloric acid, which plays an important part in beginning the digestion of proteins, and has some effect also in splitting cane sugar to monosaccharides. It contains also rennin, which coagulates the casein of milk. Its most important enzyme is pepsin, which, as the previous experiment has shown, splits proteins into proteoses and peptones, substances very similar to each other. These have smaller molecules than proteins, are soluble, and so are ready to be absorbed by the capillaries of the intestinal wall. It will be remembered that a similar kind of change was caused by the action of ptyalin and amylopsin on starch, and of lipases on fats. The action of pepsin is also catalytic and the process is essentially hydrolysis. Another ferment called trypsin, found in the pancreatic juice in the intestines, has this same power, and it is now understood that the trypsin is even more necessary than the pepsin, since the food in the intestine is exposed more intimately and for a longer time to its action. The digestion of protein is carried still further, with some of it at least, by erepsin of the intestinal juice. This enzyme forms from peptones amino-acids, the simplest cleavage products of protein food.

It remains to discuss the chemical composition of proteins and their use in the body. They are the most complex of the nutrients, and the most essential. They make up a large group, with many subdivisions, but are all now considered to be essentially anhydrides of a group of simpler nitrogenous substances called amino-acids. It has been calculated that the molecular weight of some proteins must be over 16,000, which shows their enormous complexity. They contain the elements carbon, oxygen and hydrogen, small quantities of sulphur, iron and phosphorus, and in addition nitrogen. Since protoplasm, or the living substance of cells, is a complex compound of nitrogen, and since waste products containing nitrogen are constantly being excreted from the body, chiefly in the form of urea (CON_2H_4) from the kidneys, so nitrogenous foods are evidently a necessity. It is not possible to discuss with any approach to completeness the rôle of nitrogen-containing foods. It is known that they are received from the digestive tract by the blood as peptones, or as amino-acids, that as they are so received they are changed again into proteins, that they are taken up by the cells of the tissues and by them built up into protoplasm. They are indispensable as food not only because they serve to produce energy by their oxidation, as do carbohydrates and fats, and to aid in the storage of glycogen and fat, but because they are probably the only material out of which the substance of the active tissues, muscle, nerve, gland, are made. These tissues in performing their work are constantly breaking down or disintegrating, and must therefore be as constantly repaired. Proteins are the only nutrients able to do this. A person might exist, very uncomfortably and not at maximum efficiency it is true, on a diet purely protein; but he would in time starve to death on a diet, however abundant, that contained only carbohydrates and fats.

A proper proportion of protein food in the diet is a very important matter. When carbohydrates and fats are oxidized, carbon dioxide and water are the only wastes produced. These two substances are at once carried to lungs, skin and kidneys by the blood, and excreted without any further work on the part of any organ. When proteins are oxidized in the body, the products formed are not only carbon dioxide and water, but urea, creatin and uric acid, substances produced only after rather complicated reactions in the liver and kidneys. This is one chief reason for not unduly increasing the amount of protein eaten. Too much protein imposes an unfair strain on liver and kidneys, and may mean the accumulation in the body of such wastes as uric acid, which is very harmful. Protein food is also especially subject to bacterial decomposition in the digestive tract. Such decomposition forms substances which are always objectionable and if absorbed into the blood at best lower the tone and vitality of the tissues and may produce definite cases of poisoning. Too little protein, on the other hand, means that the body lacks those nitrogenous compounds which are absolutely necessary for the upbuilding of the tissues. It is estimated that for growing children the diet should contain one part of protein to every four of carbohydrate and fat, and that for adults it should contain one part to every five or six of carbohydrate and fat. Any decided increase or decrease in this proportion is undesirable.

Within these limits various considerations govern the proportions of these three classes of nutrients advisable in any diet. The age and weight of the person, and whether he is leading an inactive or active life are important, as well as the degree of ease with which different people digest the different foods. There are some general statements that may be made. (It has been found that the best balanced diet contains for one day 100 grams of protein, 100 grams of fat, and 420 grams of carbohydrate. In avoirdupois weight these figures are 3½ ounces of protein, the same weight of fat, and 15 ounces of carbohydrate, equaling 22 ounces of solid oxidizable nutrients which would be contained in about 4 pounds of ordinary food. The energy produced by the oxidation of this amount of food is equivalent to 3000 Calories of heat, which would be sufficient to raise ten gallons of water from ordinary temperatures to the boiling point; or put in terms of mechanical work would raise 2000 pounds, or a ton, 4500 feet high, or almost a mile.) This gives some idea of how extensive are the chemical changes which take place in the body every day.

The abridged outline of digestion on page 172 will help in summarizing the process.

QUESTIONS

1. In making cocoa, why is milk added after boiling the cocoa, and not before? Why is it well to beat the cocoa before serving?
2. Considering the plant and the animal as chemical laboratories, in which do you think that the reactions are mainly endothermic? In which mainly exothermic? Explain.
3. Review the subject of digestion by writing a short paper, discussing the chemical changes that take place in the different nutrients in their proper order of digestion in mouth, stomach and intestines.

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DIGESTIVE TRACT	*Gland*	*Juice*	*Enzyme*	*Action*	*Condition*
Mouth	Salivary	Saliva	Ptyalin	Starch to sugar	Alkaline
Stomach	Gastric	Gastric	Pepsin	Protein to proteoses and peptones	Acid
			Rennin	Casein to curd	
			A lipase	Splits emulsified fats	
Small Intestine	Pancreas	Pancreatic	Amylopsin	Starch to sugar	Alkaline
			Trypsin	Protein to proteoses and peptones	
			Steapsin	Splits fats	
	Liver	Bile		General aid	
	Intestinal	Intestinal	Several	Peptones to amino-acids Sugars to simpler sugars Furnishes Na_2CO_3	

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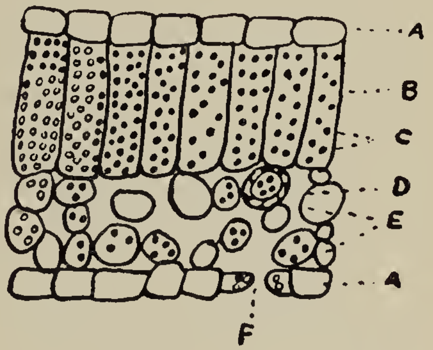
CHAPTER IX

CARBON AND NITROGEN CYCLES

THE arrangement of this text implies a wider separation between inorganic and organic compounds than actually exists. The inorganic and organic worlds are of course mutually dependent. If you think of human beings with their primary needs of air, food and water this dependence is evident. Our food is mainly from the vegetable or from the lower animal world. The animals are for a large part dependent on vegetable food. The vegetable kingdom derives its food from the soil, air and water, which are in turn made up of elements, oxides and salts — inorganic substances. All elements occurring in living bodies may be traced in cycles from this organic to inorganic nature, and back again. Especially striking are the cycles of carbon and nitrogen.

Carbon, one of the commonest of elements, is an important link which holds together the plant and animal world and makes one of the most interesting phenomena in the living world. It is interesting partly because man's ingenuity has not yet solved the problem, and therefore has not been able to imitate it or duplicate it in the laboratory. What we cannot produce ourselves is apt to be held in awe.

In the chapter on carbon compounds, carbon dioxide was found to be given off in the exhaled breath. This carbon dioxide results from the oxidation of the carbon in the food, and is given off by all animals, thus filling the air with carbon dioxide. Since this gas is thrown off by the animal body it is no longer of any use to that body. With all the animals on the earth the atmosphere would soon be heavy with this gas, so useless to animals. Analyses of the air however show a very small percent of it. Nature has very wisely provided for its consumption. People who have studied the life histories of plants find that the ordinary plant gets its food in two ways. One kind of food the plant takes in through its root; namely, salts in the soil which are dissolved in rain water. Another kind of food the plant takes in through little openings in the leaf called stomata. The food so consumed is the carbon dioxide of the air.



Cross Section of Leaf

Greatly magnified

a. Epidermal tissue
b. Palisade cells
c. Chloroplasts, containing chlorophyll, the green coloring matter
d. Vein
e. Spongy tissue
f. Stoma, space through which carbon dioxide enters

In the plant body this carbon dioxide together with water undergoes a tremendous change, a change so complex that man is still ignorant of the process. He knows the initial substances and the end products, but just how one substance is formed from the others is still a mystery. Botanists tell us that the green coloring matter in the leaf, chlorophyll, together with energy of the sunlight, enables the plant to turn this carbon dioxide and water over into starch and sugar.

$$6CO_2 + 5H_2O \rightarrow C_6H_{10}O_5 + 6O_2.$$

The simple molecules CO_2 and H_2O form a very complex molecule $C_6H_{10}O_5$ or some multiple of this formula. Oxygen is sent out as a by-product by the plant in the daytime into the air, thus making the air better for animals and for people. This difficult task is done by the plant with its green leaves in the presence of sunlight. The green plant is truly a manufacturing concern. This starch the plant may store up in large quantities for future use in tubers, as in the case of potatoes; in its roots, such as the carrot and turnip, which are common starchy "vegetables"; or in smaller quantities in stems and leaves, as in celery, cabbage and so forth. Or the plant may use this starch in the manufacture of sugars, like the sap of the maple tree; of fats and oils, as is found in many fruits of the nut variety; and of proteins, which make the living substance or protoplasm of plant cells. All of these substances we and the lower animals eat as foods, and in the animal body, by processes of digestion already studied, such substances are made ready to be used by the cells. Finally by uniting with the oxygen breathed in through the lungs such substances are decomposed and oxidized into simpler waste substances, chiefly urea, carbon dioxide and water, which are eliminated by the excretory organs — the kidneys, lungs, skin and intestines. Some carbon is removed in combination in the urea, but most of it reaches the air in carbon dioxide, and so the cycle is again and again repeated.

In much the same way as the element carbon, the element nitrogen makes its journey from inorganic to organic world and back again. Some of the plant foods found in the soil are nitrates and ammonium salts, all of which contain nitrogen and are inorganic in nature. Taken into the root of the plant these nitrates and ammonium compounds are, with the help of the elements in carbon dioxide and water, built up by some life process into the more complex proteins and protoplasm. Just as the complex starch is made from simple carbon dioxide and water so the complex proteins and protoplasm are made out of these simpler nitrogen compounds. Again, we know the initial substances and the end products, but just what happens in between is yet unknown.

These complex nitrogenous substances, as food in the animal body, are attacked by certain digestive juices which render them fit for absorption into the blood supply. Oxygen from the lungs also gets into the blood stream so that when carried to the tissues a process of oxidation goes on in these tissues. During this process the nitrogenous substances do their special work for the body, building up new tissue or repairing old tissue, at the same time forming certain by-products which are waste products for the body and must be eliminated. Water and carbon dioxide are again made as waste products, but the most important waste product given off by the work of nitrogenous material in the body is urea $(NH_2)_2CO$, a crystalline substance given off entirely in solution by the kidneys in urine, and to a much smaller extent by the skin in perspiration.

This urea from animal bodies finds its way into the soil where, undergoing decomposition, ammonia gas and nitrates are again made, being given into the air or left in the soil to be used as plant food. This change repeated again and again makes the nitrogen cycle as complete as the carbon cycle. With the elements carbon and nitrogen in their circular movements are combined the elements oxygen, hydrogen, sulphur, phosphorus and a few others, but these others do not produce so distinctive a cycle as carbon and nitrogen because they are not responsible for so definite a group of compounds as the carbohydrates or the proteins.

Not every atom of carbon or nitrogen completes this round of existence. Probably this is more true for the nitrogen atom than for the carbon. Nitrogenous wastes from animal bodies in the process of decomposition produce ammonia gas which may escape into the air and so be lost, for in general plants can use only the gas carbon dioxide. Evidence of this loss of ammonia is frequently found in barnyards, where the strong odor of ammonia in the air indicates the decomposition of nitrogenous substances and the escaping of the ammonia into the air. Such a loss means in time a lack of nitrogen compounds in the soil. This lack is very serious, because both the plant and the animal must have the nitrogen in order to build up body tissue. When soil lacks nitrogen compounds man has to introduce it in some way. This used to be done by using the excrements of the farm animals. This material when rotted puts into the soil the nitrogen compounds in the form of ammonium compounds and nitrates. With the increase of land cultivation and the decrease in cattle breeding such material is hard to obtain, and so chemical fertilizers have largely replaced this so-called natural fertilizer.

The most common of these fertilizers is Chili saltpetre, a sodium nitrate found in large quantities in Chili and thought to have been formed from the decomposition of animal remains of long ago. Phosphates, found naturally in the South, or made from animal bones, are another class of these chemical fertilizers in common use. Different salts, such as sulphates, ammonium salts and other nitrates, are now made to help out the supply of these naturally occurring compounds.

The supply of natural fertilizing material containing nitrogen is limited, and yet nitrogen is constantly escaping from this cycle and being added to the supply of free nitrogen which makes up so large a part of the atmosphere. This nitrogen supply is kept constant by the decomposition of organic material, since such material is attacked by micro-organisms which produce nitrogen as the simplest end product. The nitrogen found in the air, though easily obtainable, is not so easily usable for chemical reactions. Neither animals nor plants can use it ordinarily. It is an inert gas under ordinary conditions, so that to convert it into compounds extraordinary conditions are necessary. It is evident that there must be ways of "fixing" it or causing it to combine with other elements so that it may be of use to the organic world. These ways may be classified as artificial and natural.

The artificial ways of utilizing the nitrogen of the air are first, the liquefying of the air which separates out the nitrogen which is used to combine with other substances; and secondly, by combining the nitrogen and oxygen of the air by the action of a heavy electric current. Recent developments in the production of high temperatures by an electric arc furnish conditions which make the conversion of atmospheric nitrogen into compounds not only possible but produce these most desirable substances in quantities at a low cost. This fixing of nitrogen is one of the most important achievements of modern science, and especially when war conditions demand the use of such large quantities of nitrogen compounds for the manufacture of high explosives, it represents a large commercial activity.

Calcium carbide, which is familiar to many people as the source of acetylene, has the peculiar property of taking up nitrogen to form a cyanamide if the action is carried on in the electric furnace. Calcium carbide is acheap substance, and when heavy electric currents can be procured the process of making the cyanamide is not expensive.

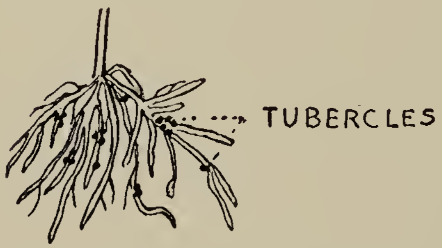
$$CaC_2 + N_2 \rightarrow CaCN_2 + C.$$

The cyanamide can be used as fertilizer under certain conditions and is of value both for the nitrogen and calcium content.

Another method combines the gases nitrogen and oxygen directly as they exist in the air. A large electric spark is passed through air which produces first the compound nitric oxide, then nitrogen peroxide, which treated with water forms nitric acid. The acid is then treated with lime, which produces calcium nitrate, in which form the nitrogen is " fixed " to be put on the market.

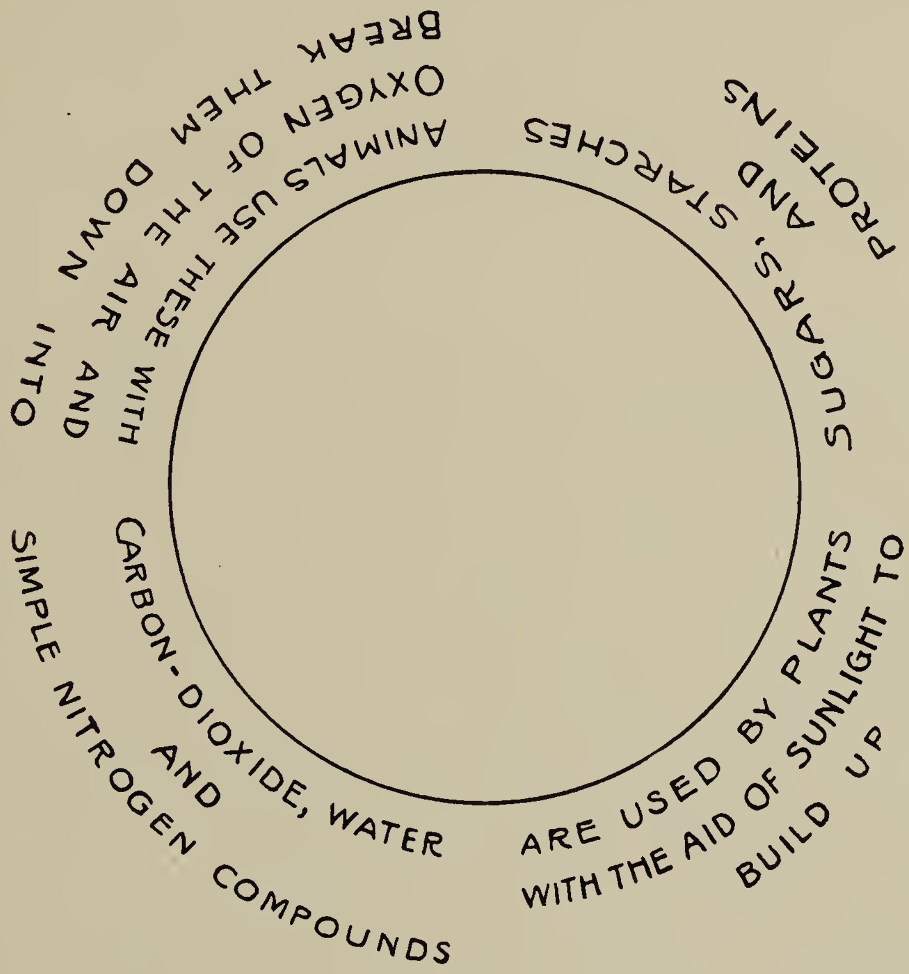
Both of these methods are used in Europe at the present time, especially in Sweden where water power is cheap. In this country the same general methods have been tried, especially at Niagara Falls, but as yet the processes used are not on such a paying basis as those of Europe.

The natural method of using atmospheric nitrogen takes us back again to the plant kingdom, and illustrates again the interdependence of the organic and inorganic worlds. In general, green plants use as food material the carbon dioxide of the air; but there are some plants like clover, peas and beans, and other leguminous plants, which have the extraordinary power of using directly the nitrogen of the air. They do not however absorb it through the stomata of the leaves as they do carbon dioxide. Leguminous plants bear on their roots small white knobs which are called tubercles. These tubercles contain bacteria which are capable of building gaseous nitrogen, together with the help of other substances, into nitrites, and nitrates in the roots. These compounds are then built up into proteins by the plant's activities. And again the simple inorganic substance is transformed by processes in the living body into complex organic substances. The fact that such plants as clover and peas manufacture nitrates makes them very valuable



Root of clover, showing tubercles containing bacteria which use atmospheric nitrogen.

as fertilizers. Such plants are often plowed under the soil, thus giving the organic and inorganic material to a worn-out soil, and are then spoken of as a " green " fertilizer. When several crops are planted in the same soil one crop ought to be of the tubercle-bearing root variety in order to keep the soil rich enough in nitrogenous plant food.



With both carbon and nitrogen compounds it is evident that plants are the building-up organisms and animals the breaking-down organisms. Plants use elements and simple compounds to synthesize complex compounds. Animals eat these complex compounds and break them down into simpler ones and into elements. So the two great kingdoms are interdependent.

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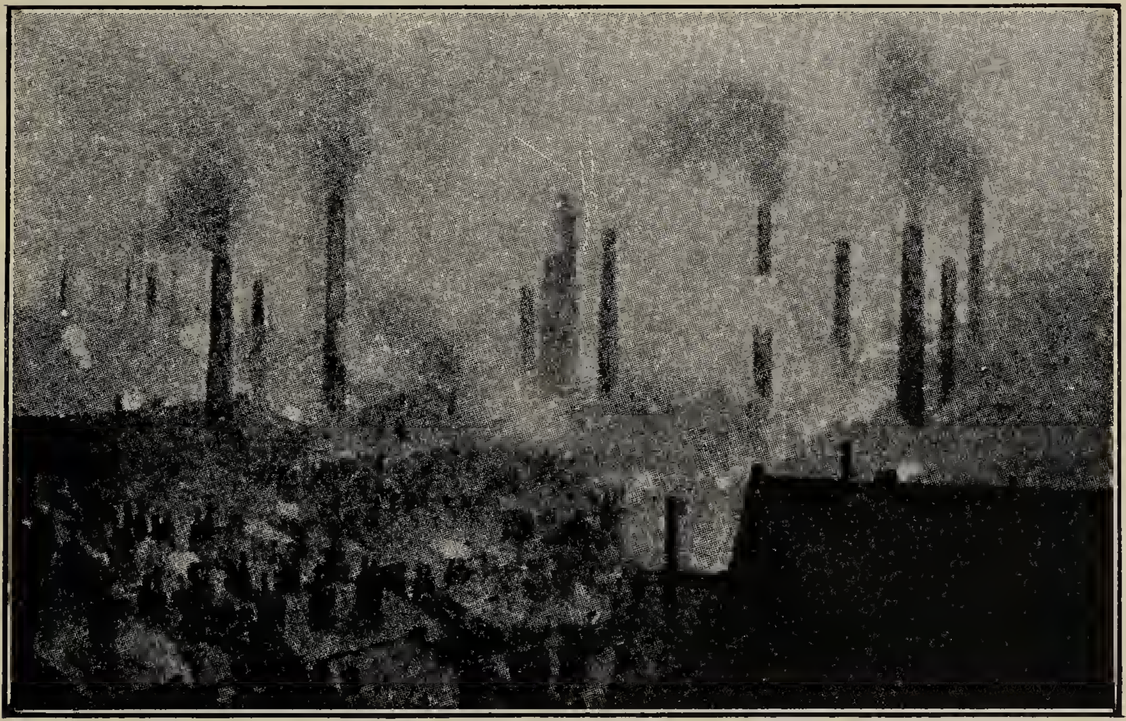
CHAPTER X

CLEANING AGENTS WITH WATER

THE increasing complexity of modern living is constantly increasing the need of cleaning methods. At the same time increased knowledge of substances tends to make this cleaning more successful in every way. The thousand and one activities of the private house together with those of the community are almost constant sources of "dirt." The business of removing this dirt offers a variety of chemical problems. Dirt is a very large term. To the bacteriologist it means one thing; to the housewife



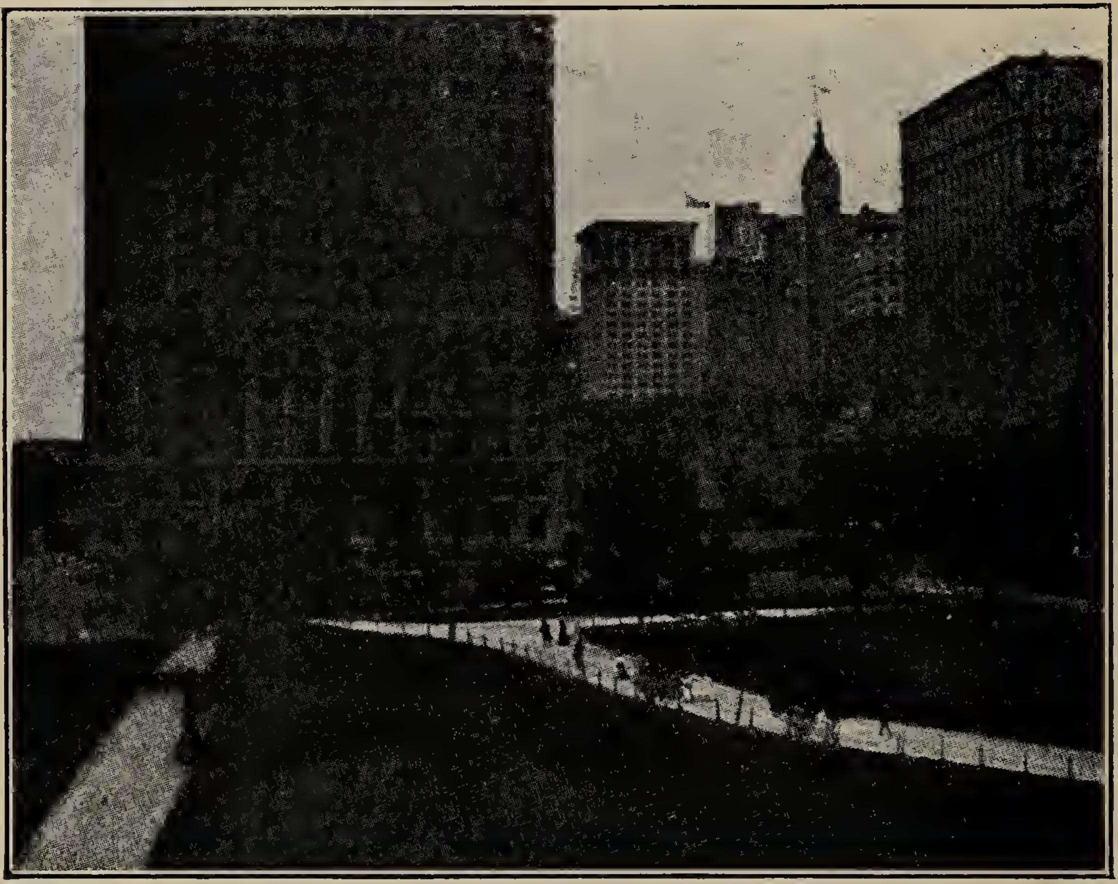
From "Municipal Chemistry"
Edited by Baskerville, published by McGraw-Hill Book Co.

Leeds, England, overlooking Kirstall Road

This scene may be observed any hour of the day in Leeds.

it means another; and to the child making mud pies in the back yard it has a very different meaning. The bacteriologist usually means that the presence of germs indicates dirt, while the housewife means dust, or grease spots, or tarnish perhaps, or spots and stains on clothing.

Dirt has been defined as "matter out of place." For our discussion of the chemistry of cleaning perhaps that is a good statement. The dirt we mean is matter out of place and includes dust and grease; spots which appear on metallic surfaces as they are exposed to the air; spots made on fabrics by fruits, foods or other such substances.



From "Municipal Chemistry"
Edited by Baskerville, published by McGraw-Hill Book Co.

A View of Down-Town Skyscrapers from Battery Park

Showing very little smoke issuing from these tall buildings, each of which operates its own power plant.

As long as people live in crowded communities, and as long as factories continue to pour out into the air fumes and smoke, much material is let loose which is finally deposited in our houses as " dirt." This particular kind of dirt is truly matter out of place in many cases and furnished our grandmothers a fundamental cause for that tremendous upheaval known as spring housecleaning. This kind of dirt which is not at all snobbish gets onto and into everything in the house, furnishings and hangings, clothing, wall paper and so on. Nothing in a house is free from it.

Any form of dirt, wherever it may be, has to be removed by the use of some kind of cleaning substance.

The most common cleansing agents are water, soap, washing soda, borax, hydrochloric acid, oxalic acid, pumice, sand, chalk, rouge, diatomaceous sand, kerosene, turpentine, chloroform, ether, benzene and naphtha, and various bleaching agents. Water is the most common and natural cleansing agent because it is the most common solvent. Occasionally it is used by itself but most frequently it is used with soap, or borax, or ammonia or washing soda. The action of water with these various substances furnishes much chemical discussion. So we begin with the preparation of soap.

EXPERIMENT 31

Purpose.

To show the action of strong alkali with oil.

Directions.

To a small amount of very concentrated caustic soda solution in a test-tube add a drop or two of olive oil. Shake *gently,* noting whether the oil goes into solution or not. Allow it to stand some time.

Questions. Does the oil go into solution? If not, describe appearance. Does the appearance of contents of test-tube change when allowed to stand?

NOTE. An emulsion is obtained here. An emulsion is the intimate mixture of an oily substance with water.

EXPERIMENT 32. (Class Experiment.)

Purpose.

To make soap — Cold Process.

Directions.

a. Take half these quantities:
1 can Babbitt's Lye.
1 qt. cold water.

Slowly add water to lye, stirring with long-handled wooden spoon. Stir till cold.

6 lbs. melted and strained grease, lukewarm.

Dissolve in $\frac{1}{2}$ cup cold water, 4 tbsp. borax, $\frac{1}{4}$ cup ammonia, 4 tbsp. sugar, 1 tsp. salt. (Sugar and salt make a lather.)

Pour lye into grease very slowly, add other mixture, stir or beat until thick. Pour into a pan lined with cheesecloth. Let it stand until it hardens.

b. Testing for free fat: Few shavings of soap made in *a* shaken up with cold gasoline; filtered; evaporate filtrate over warm water. Greasy residue indicates unsaponified fat.

Testing for free alkali: Few shavings of soap made in *a* shaken up with warm 95 per cent alcohol; filtered and drop of phenolphthalein added. Red color indicates free alkali.

When fats and oils are boiled with strong alkalies like sodium or potassium hydroxide and sodium and potassium carbonates the solid which is formed is called "soap." The fat or oil is a substance made by either a plant or animal and is chemically an organic salt (ethereal salt or ester), just as a chloride is an inorganic salt. Which particular salt any one fat will make is determined by the fatty acid to which it is related. The three most common salts are tristearin and tripalmitin, fats which are solid at ordinary temperatures, and triolein, a fat which is liquid at ordinary temperatures (oils). The acids to which these are related are palmitic, stearic and oleic acid.

When chemical action occurs between these fats and an alkaline hydroxide, the same principle is involved as in Experiment 7 in Part I. There is formed a base glycerol or glycerine and another salt, sodium or potassium stearate, or palmitate or oleate, as the case may be. This insoluble palmitate or stearate or oleate is what we know as "soap." This process is known as saponification. Soap is the cleansing agent always used with water. In the earlier part of the work water was found to be very necessary to any chemical action. And here again the same condition holds.

EXPERIMENT 33

Purpose.

To study the action of soap and water.

Directions.

a. Shave a little ivory soap (piece as large as dime), and dissolve the shavings in a test-tube half full of alcohol. Add a drop or two of phenolphthalein solution. Shake, and note any change. (A good soap should give no change.)

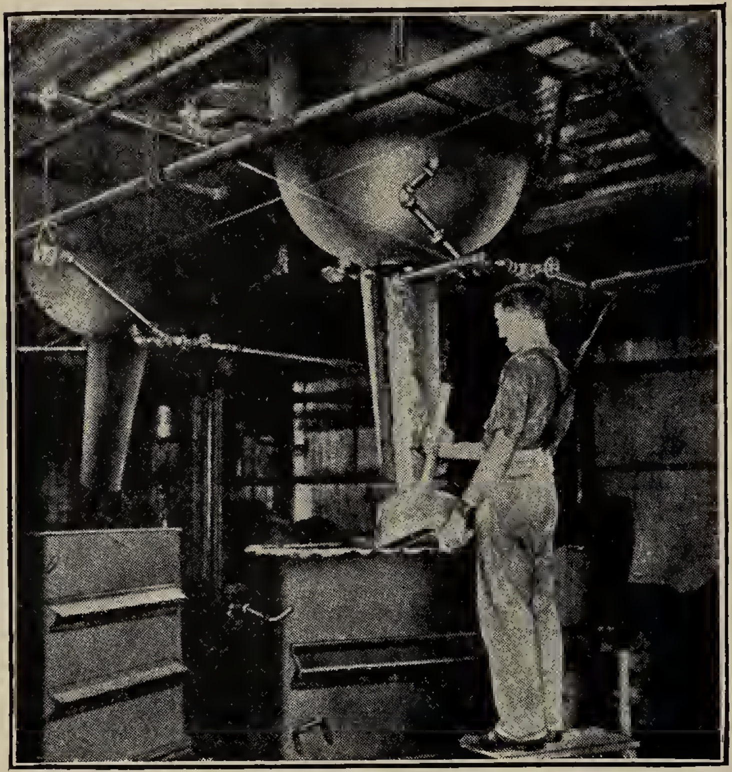
b. Put same amount of the shavings into a test-tube half full of water to dissolve. Add as much phenolphthalein solution as in *a*. Compare results.

NOTE. Phenolphthalein is an indicator like litmus, and gives a magenta pink color with an alkaline substance.

Questions. What is the difference between a water solution and an alcohol solution of soap? The water has what action? What is the alkaline substance produced when soap (sodium stearate) is dissolved in water?

NOTE. Refer to chapter on ions.

The cleansing action of soap depends upon these characteristics brought out in the last three experiments. When soap is dissolved in water evidently some sodium hydroxide is formed which gives the alkaline reaction with the indicator. Whenever soap and water are used together to remove the ordinary greases, fats or oils, the soap forms hydroxide with water and this probably unites with greases forming a soap again. Soap and water also have an emulsifying action on fats and oils, so that there is a double help in the removal of foreign matter in the form of grease. The original soap has an emulsifying action. At the same time new soap is formed which in its turn helps emulsify, and so there is a continuous "performance." Both the emulsifying and more-soap making quality of soap depend very much upon the water present, both for the chemical action involved and the mechanical removal by solution of the products of the chemical action.



A Crutching Machine

Showing how the ingredients are mixed to make soap.

Substances are often put into soaps during the process of making to give special properties. Some of the common ones are carbolic acid, oatmeal and such substances as tar.

Liquid soaps are made by the same process of saponifying fats, but with the addition of alcohol. Usually a liquid soap is purer than some solid soaps, for there is not so good a chance to add fillers or detergents on account of their lack of solubility in alcohol.

Scouring soaps usually contain a small amount of soap and a large amount of abrasive material, such as sand, pumice, whiting or ground slate. Some contain sodium carbonate.

Soap powders are made by melting sodium carbonate and adding soap, then grinding the hardened mass to powder. They may contain little soap or much soap, and much water.

A good laundry soap must mix with water easily, must contain no unsaponified fat, and very little free alkali. For toilet soaps or laundry soaps for woolens there should be no free alkali. Since the fats used in the commercial preparation of soaps are mixtures of the glycerides it is no easy matter to get a soap free from fat or alkali. Free fat of course defeats the purpose of the soap, and free alkali if present in large quantities has a bad effect on the hands, on some textiles, on painted or varnished surfaces, and on aluminum ware.

Toilet soaps that are colored, or have a strong perfume, are often to be avoided. The colors and perfumes are many times put in to conceal the poor quality of the soap.

Such soaps as Castile soap and Ivory soap are in the long run most satisfactory. Transparent soap is made by dissolving the soap in alcohol and then allowing the alcohol to evaporate. It is considered a purer soap by some people, but that is not necessarily true. Marine soap is a soap made of palm nut or cocoanut oil. It contains much water and forms a lather with sea water. Soaps sometimes, besides containing unused alkali from the original substance, contain other alkalies which perhaps help in the cleansing action but which are really adulterants. These are often used to lessen the cost of production, so that altogether they are not to be desired. The most common of these are sodium and potassium carbonates, borax, rosin and petroleum products like kerosene and naphtha.

"Fillers" are sometimes used to make soap heavier and to cheapen the soap. They have no cleansing action usually, and are of course harmful in the use of such soaps because they are insoluble in water. Some of the most common "fillers" are sodium, potassium, calcium and barium sulphates, infusorial earth, chalk and starch.

Ammonia, borax and washing soda are next to soap most frequently used with water as cleansing agents. An experiment in making ammonia is given on page 249. The water solution known as ammonium hydroxide is the common form in which it is used. Both sodium borate, commonly called borax, and washing soda (sodium carbonate) are white crystalline salts soluble in water. The following experiment may help in showing the reason for the use of these as cleansing agents.

EXPERIMENT 34

Purpose.

To study the action of household ammonia, washing soda, borax and soap, on fats.

Directions.

Fill a test-tube one-third full of each of the following solutions: ammonia water, washing soda, borax, borax and soap. Test each with litmus. Put a few drops of olive oil into each of the test-tubes. Shake gently.

Questions. The litmus test indicates the presence of what substances in these tubes? In which test-tube does the oil most completely disappear?

Acids like hydrochloric or oxalic may be used for cleaning purposes. They are always used in water solution, and their chemical activity as acids is so familiar from previous work that it needs little comment here.

Another class of substances nearly always used with water are the scouring materials. These scourers are practically all insoluble in water, but in order for them to do their work well and to come in contact with all parts of the article to be cleaned they must be mixed with water. This mixing increases the area of contact and keeps the substance in place. Their action is however practically all mechanical.

EXPERIMENT 35

Purpose.

To test for insoluble residues in scouring soaps and soap powders.

Directions.

a. Procure samples of Sapolio, Bon Ami, or any scouring soap. Dissolve a little (about ½ teaspoonful) of each of the samples in an Erlenmeyer flask with about 50 c. c. of water. Heat gently. Filter off the insoluble part. Compare the original bulk with that of the insoluble part. Test the filtrate with litmus. Feel the residue on paper to determine its coarseness. To the residue on filter paper in the funnel add hydrochloric acid. Effervescence indicates the presence of insoluble carbonates. Add acid until there is no further action. If there is a residue it is probably silica, SiO_2.

b. Procure samples of Pearline, Gold Dust or any soap powder.

Dissolve a little as in *a*, then proceed as in *a*.

Record the results.

QUESTIONS

1. A familiar expression is that soap "cuts" grease. Chemically, what does that mean?
2. Why cannot a soap like Fels-Naphtha, which contains a petroleum product, be used with hot water?

3. Would you prefer a colored or uncolored soap for toilet purposes? Scented or unscented? Transparent or opaque?

4. On the whole, is transparent soap more or less expensive than an opaque one? (Weight is the determining factor.)

5. What is the advantage of liquid soap?

6. Why is it better to use hot water than cold water with soap?

7. Is ammonia water, soap and water a good combination? Of what value may the ammonia water be?

8. Which is the more vigorous acid, hydrochloric or oxalic?

9. Why is Dutch Cleanser such a satisfactory cleaning agent from the chemical point of view?

10. Is it economical to buy crystallized sodium carbonate?

11. Discuss the use of Sapolio, Bon Ami, etc., based on your own experimenting.

12. Find out the cost of chalk, whiting, or any abrasive per pound and compare it with cost of polishes and scouring soaps.

13. What kind of scourer should be used to clean the bottom of a kettle?

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Disposal of Wastes. Baskerville. Municipal Chemistry, Chaps. XIV, XVIII, XXI.

CHAPTER XI

WATER IN CLEANING

WATER used for cleansing purposes may come from a variety of sources; from the air as rain water, from surface supplies as ponds, rivers and lakes, or from underground sources such as springs or artesian wells. Rain water is of course pure or soft, but not always available in these days of living in crowded cities. The purer the water the better solvent action it has, and therefore the better cleanser it is. Surface water has many chances to get mineral matter dissolved in it, as does also underground water. Such water is hard water, and has less solvent action than soft. The particular kind of material which makes water hard will be determined by the rock material over which or through which it flows.

EXPERIMENT 36

Purpose.

To study temporary hardness of water.

Directions.

a. To half a test-tube of distilled (soft) water add 10 c. c. of a soap solution previously prepared. Shake vigorously, noting any change.

b. Fill a test-tube two-thirds full of limewater. Pass carbon dioxide into it until solution is clear.

NOTE. This is hard water.

c. Divide the solution made in *b* into halves. To one half add 10 c. c. soap solution. Compare with *a*.

d. Boil gently the second half. Filter and boil again, filtering again if necessary. When repeated boiling makes no change add the same amount of soap solution as in *c.* Shake and compare results with the unboiled one (*c*).

Questions. What do you see when soap is dissolved in soft water? In hard water? Why does the precipitate disappear when more carbon dioxide is added? What is condition of the water now? Why does it reappear when water is boiled? What is now the condition of the water?

NOTE. Soap solution is made by dissolving 10 gms. of Castile soap scraped into shavings in a liter of alcohol diluted with one-third water. Filter.

EXPERIMENT 37

Purpose.

To study permanent hardness of water.

Directions.

Make some hard water by dissolving calcium sulphate in water, making about 80 c. c. of the solution. Filter. Divide into four portions of about 20 c. c. each.

a. Add 10 c. c. soap solution to one part. Shake vigorously. Compare with Experiment 36 *c.*

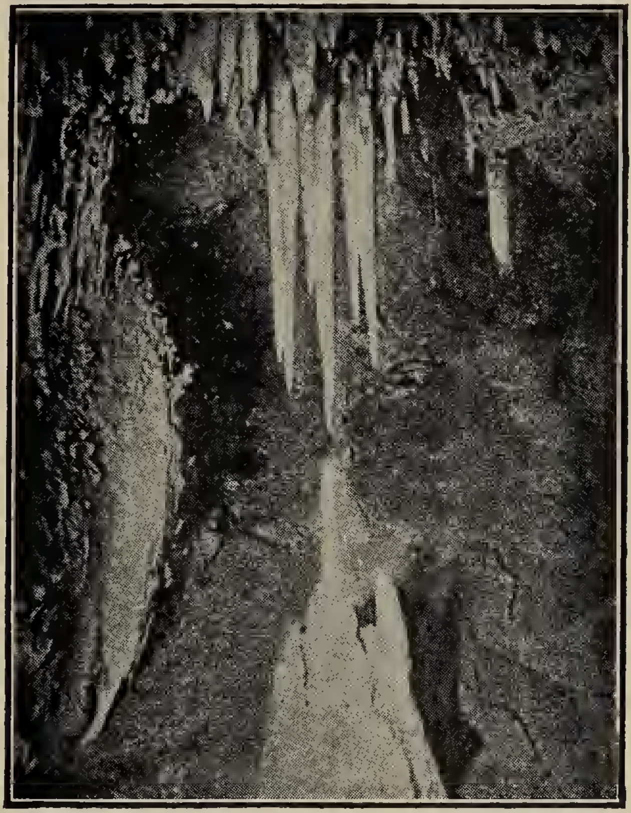
b. Boil a second portion. Add 10 c. c. soap solution. Compare results with Experiment 36 *d.*

c. To a third portion add a solution of sodium carbonate. Filter and test filtrate with sodium carbonate. When no further precipitation occurs test the hardness of filtrate with 10 c. c. soap solution.

d. To the fourth portion add a solution of borax. Is there evidence of chemical action?

Questions. What effect has boiling on this kind of hard water? In *c* what substances are formed? Which one is left in the water (filtrate)? Would it do harm if present in excess? Can you see why calcium sulphate is said to make water permanently hard?

From the foregoing experiments it is easy to understand what is meant by "hard water." Water with any dissolved mineral matter in it feels hard to the touch, and often has a roughening effect on the skin. If the mineral matter is calcium or magnesium carbonate the water is said to be temporarily hard. This hardness is partly removed by boiling. The presence of excess carbon dioxide makes the carbonate soluble, or forms according to some authorities an acid calcium or magnesium carbonate which is soluble. Boiling in the first case removes the excess of gas, and so precipitates the carbonate which may be filtered out. In the second case heat decomposes the unstable acid salt into the normal salt which precipitates out and may be filtered out.



Stalactites (above)

Formed by the dripping of water containing calcium carbonate in solution, which leaves the solid in the form of an icicle as the water evaporates. In the same way stalagmites are built up on the floor of cavern as water drops from the roof.

Sometimes this kind of hard water is treated with lime, but it is not so successful a method except on a large scale. When the mineral content is calcium or magnesium sulphate or chloride the water is said to be permanently hard because the impurity cannot be removed without chemical action. The use of sodium carbonate (soda ash) or borax does not produce a perfect water because some of the softeners are left in the water, but there is in both cases, however, an alkaline reaction which helps in the laundering process.

Since part of the cleansing action of water depends on its solvent power a hard water does not meet this purpose fully. From the work in a preceding chapter it is easy to see that a hard water cannot dissolve as much as a water free from solutes. The use of soap with water is greatly interfered with when the water is hard. Much soap is wasted in overcoming the hardness, and the precipitate produced by this action is extremely unpleasant and inconvenient.

QUESTIONS

1. What would be the result of long continued boiling of hard water in a tea kettle?
2. Why is rain water a good kind of water for laundry and toilet purposes?
3. Why is borax or washing soda often added to water even when soap is used?
4. How could water be tested in one's own home for hardness?
5. Why does not boiling a permanently hard water free it from its impurities?
6. If water contains mineral matter in it, which is better to boil it in, an iron or aluminum kettle?

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Water and its Purification.
    Bailey. Sanitary and Applied Chemistry, Chaps. V, VI.
    James C. Philip. Achievements of Chemical Science, Chap. II.

CHAPTER XII

CLEANING AGENTS WITHOUT WATER

IN Chapter X a list was given of the substances used for cleaning purposes, and the action of such substances mixed with water one way or another has been discussed. Water cannot be used on all surfaces or on all fabrics either. Most of us are familiar with the white spot produced by water on a polished surface, or with the spot produced by the sudden shower of rain on a silk dress or a broadcloth suit. Water as a cleansing agent has its limitations. But fortunately there are other removers of dirt which do not need water as their servant. These substances are for the most part insoluble in water and are of organic origin. The most common ones are chloroform, alcohol, ether, benzene, benzol, naphtha, turpentine, kerosene and carbon tetrachloride. Some of these are petroleum products, like benzene, benzol, naphtha and kerosene. Some of them, like alcohol, are produced by the decomposition of an organic substance. Others are produced directly from a living organism like turpentine, while the rest, like chloroform and carbon tetrachloride, are synthetic preparations. The characteristic which makes all these good cleaners is shown in the following experiment.

EXPERIMENT 38

Purpose.

To study the solvent action of kerosene, chloroform, benzene, naphtha, gasoline, turpentine, carbon tetrachloride.

Directions.

a. Drop a little lard or oil into test-tubes. To each tube add respectively the above substances sufficient to fill tube one-fourth full. Shake each well and notice whether there is a solution formed in each case. Evaporate a little of test-tube contents on watch glass on radiator.

b. Try the solubility of paint in kerosene and turpentine by putting a bit of ordinary white paint into a test-tube and adding the liquids to it.

c. Try varnish in the same way.

Many of these solvents in the last experiment have the great disadvantage of being very inflammable, and so in the hands of inexperienced people become sources of great danger. Naphtha, benzene and gasoline are extremely inflammable and should be used where there is little heat, even indirect, and where the air currents are strong. A good solvent which has recently come into common use is carbon tetrachloride, a substance looking not unlike benzene but whose kindling temperature is much higher. It is therefore not inflammable and is more like kerosene, though it evaporates more rapidly. Carbon tetrachloride is the basis of the cleansing fluid known as Carbona, and is also the basis of the fire extinguishing material known as Pyrene.

QUESTIONS

1. Account for the appearance of the hands when cleaning gloves in naphtha.
2. What is “ dry cleaning”?

REFERENCES

Distillation of Petroleum. Perkin and Kipping. Organic Chemistry, pp. 67–70.

CHAPTER XIII

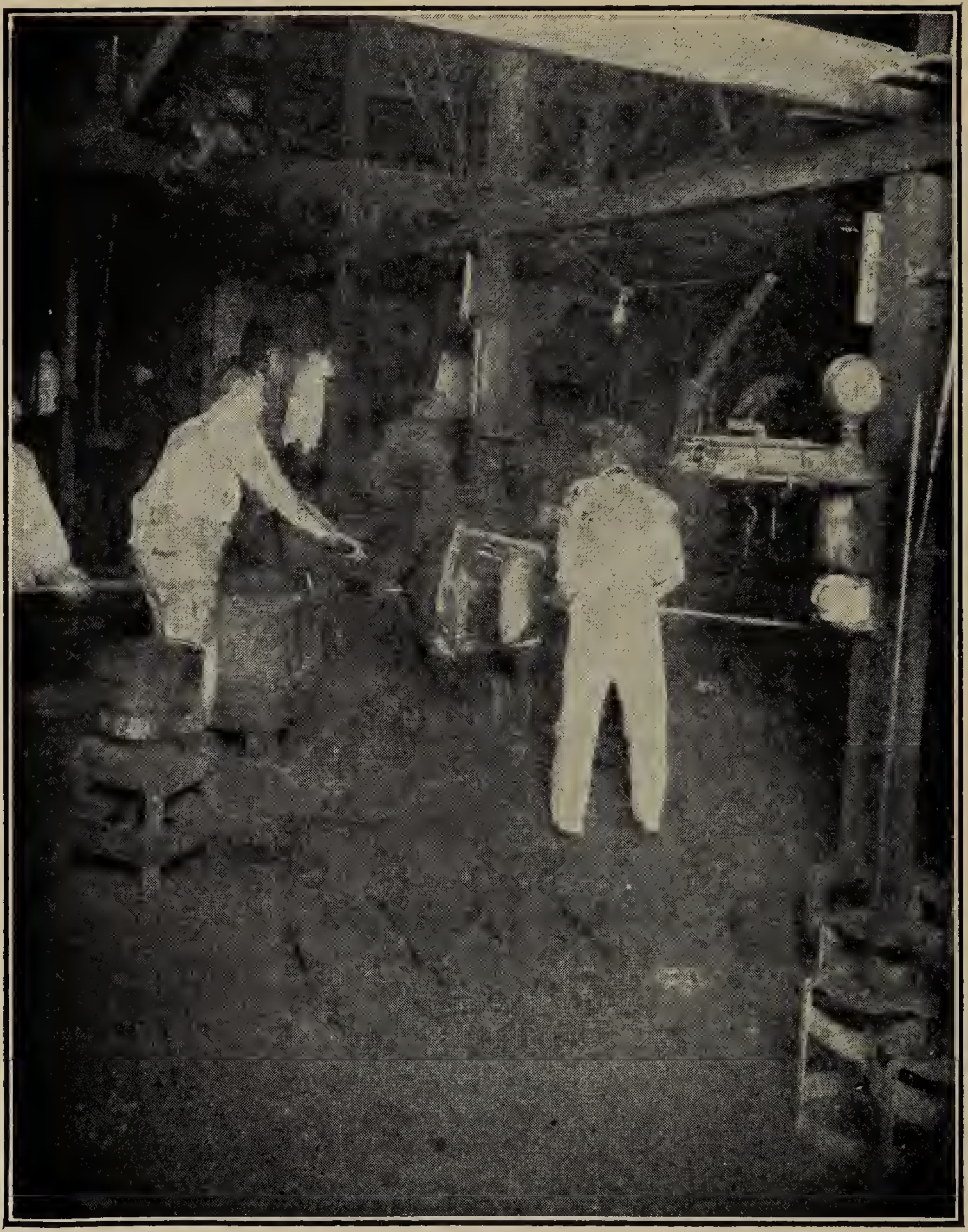
GLASS, PORCELAIN, WOODWORK AND THEIR CLEANING

MATTER so far out of its proper place as to find a resting place on the glass, porcelain and woodwork of our homes is usually removed by the very simple means of brushing, or wiping with a damp or oiled cloth. Such methods of removal involve little chemistry and therefore need no discussion here.

The cleansing of the glass and porcelain used so frequently in our dining rooms is a little more complex, and furnishes a practical application for the use of soap and water, the chemical action of which was discussed at some length in a preceding chapter.

Glass is a combination of various silicates. It is made by heating together in a furnace the following substances: white sand, potassium or sodium carbonate, calcium carbonate or some other salt. The basis is always pure white sand, and the other ingredients depend on the kind of glass wanted. Glass is a hard, brittle, transparent substance. Glass is often attacked by alkaline substances. It is difficult to make, and sometimes there is an excess of one of the ingredients. If this happens to be one of the alkaline substances the carbon dioxide and water vapor of the air attack it, making small crystals of the carbonate, which give the glass a clouded appearance.

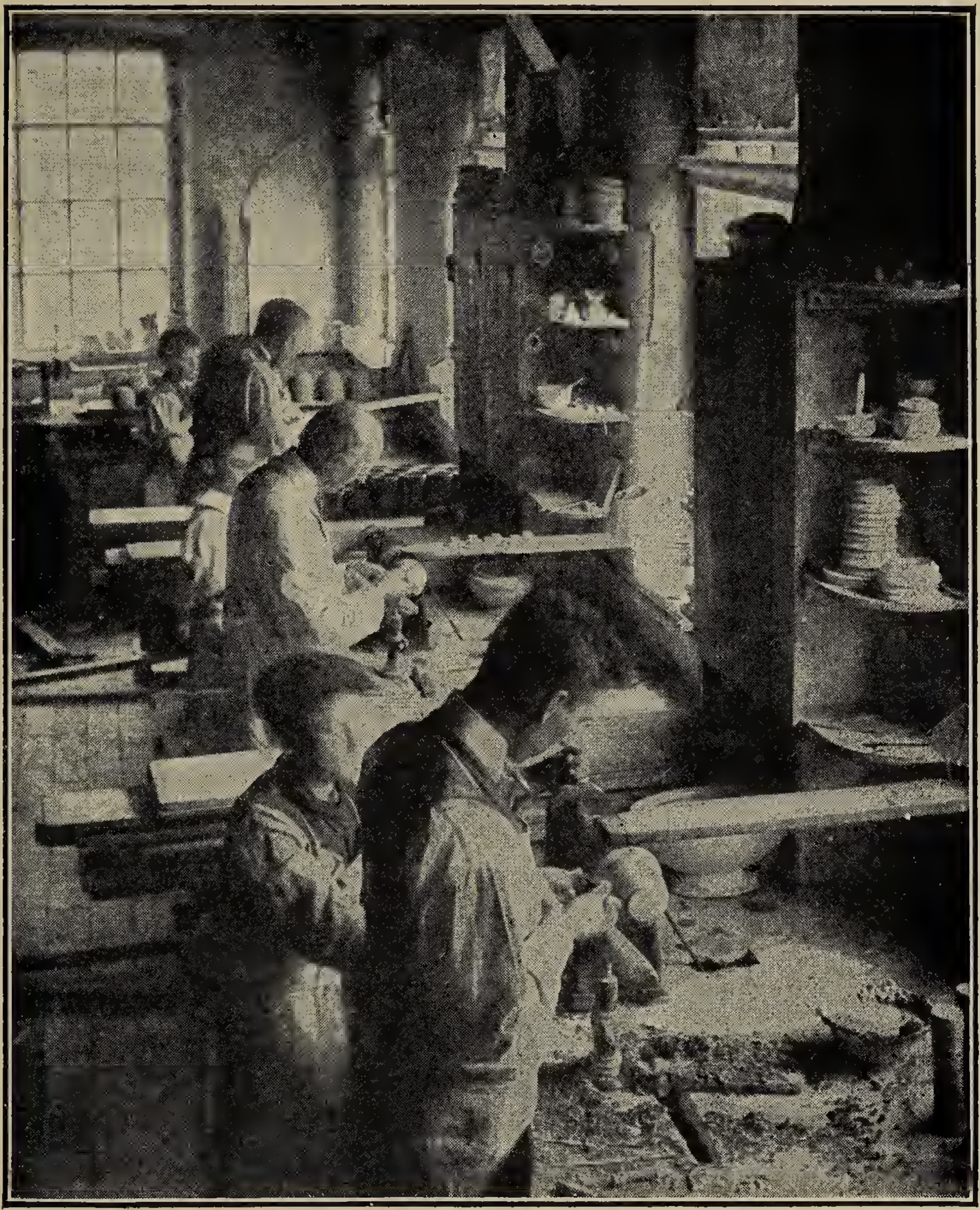
Porcelains and china are made from a fine white clay (aluminum silicate) mixed with water, made into the desired shape and then baked in an oven. The cheaper



Ⓒ Underwood & Underwood, N. Y.

Glass Bottle Making
Cutting off the neck.

grades of pottery have a porous body, but the china or more expensive grade has a non-porous body. The clay after baking is covered with a glaze and baked again. The glaze is some substance that will produce a hard, smooth, glass-like outside covering. This glaze fills up



Ⓒ Underwood & Underwood, N. Y.

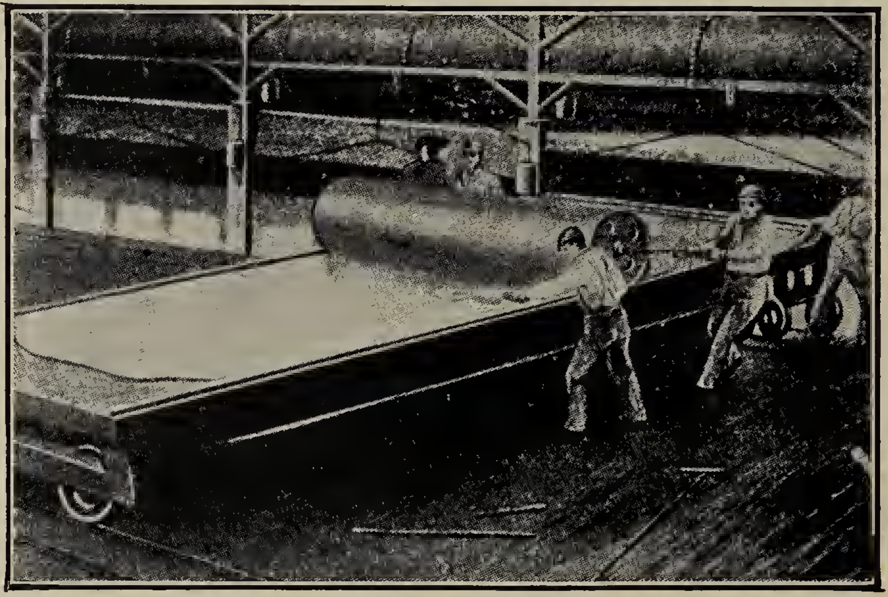
Pottery Making

Turning and trimming a cup on a lathe.

the pores and makes the vessel water tight. These glazes are not easily attacked, either physically or chemically.

The dirt found on the glass and porcelain of a house is usually of a greasy nature, or grease mixed with dust. The cleaning of chinas, porcelains and glasses requires the use of such agents as dissolve or emulsify the oily compounds. The most natural one to use is hot soapy water applied with some friction. If the article is then rinsed with clear hot water it may be wiped dry quickly enough to give a polish to the surface. Agate ware, which consists in general of an iron or steel vessel covered with a silicate glaze, may be cleaned in the same fashion. The article must not be handled too roughly or the glaze is chipped off, leaving the iron or steel exposed to the action of various corroding substances.

The dirt which collects upon the woodwork of our houses is largely of the pure dust variety or the dust mixed with oily substances. Getting rid of either of these may be accomplished by the dustless duster, which is a piece of cloth soaked in oil, or by knowing the common grease solvents. Since however there is a protective coating to all woodwork this protective coating must be respected. This coating has to be used to keep the wood smooth and to prevent decay. Soft woods are finished with paint, oil, stain, varnish or shellac, or with combinations of these; hard woods may be finished with any of


Rolling Molten Glass to make Sheets of Plate Glass

these, and in addition wax, or wax and turpentine. In most cases an abrasive will remove this coating, and we know from experiment that turpentine will remove the paint.

Painted woodwork may be cleaned with warm water and soap to which a little ammonia or kerosene has been added. Woodwork which is not painted is usually finished with a stain, varnish or shellac. A polished wood is one which has been scraped and then rubbed hard with oil. A stained wood may be cleaned by wiping with an oiled cloth, or a cloth moistened with water. If wood has a finish of varnish or shellac the problem is different. The following experiment shows clearly the effect of water on the two finishes. In general shellac should not be touched with water. Varnish may be. It is sometimes very difficult to tell one from the other, and so the safest method of cleaning is to wipe the wood with a cloth dampened with oil and turpentine, the cleansing action of which is familiar. Furniture polishes may also be used, but as they are usually the oil-turpentine combination the method mentioned above is as satisfactory. Furniture polish, in spite of its name, does little polishing — it merely removes the greasy film of dirt and helps fill up the scratches, which evens the surface.

EXPERIMENT 39

Purpose.

The cleaning of variously finished woods.

Directions.

Get several pieces of wood which have been varnished, shellacked, stained, waxed and painted respectively.

a. Try clear hot water on each.
b. Try clear warm water on each.
c. Try dilute ammonia water on each.
d. Try dilute alcohol on each.

In all cases compare the cleaned piece with the uncleaned piece.

Question. What conclusion do you reach as to the best cleansing agent for each?

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CHAPTER XIV

METALS AND THEIR CLEANING

Metals of the household, though not very many in kind or in use, furnish good objects for cleaning from the chemical point of view. The metals most commonly used are silver, copper, iron and steel, brass, tin, nickel, zinc and aluminum. Most of these are in use as such; silver and copper for water tanks, brass for water pipes, nickel for plating on pipes and stove trimmings, aluminum for cooking utensils, and iron or steel used as a foundation material for agate or tin-covered dishes.

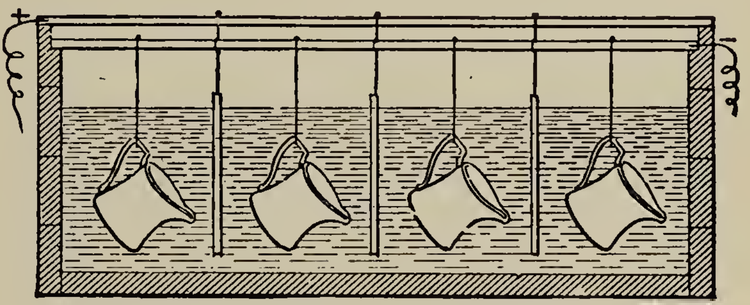


Diagram showing an Arrangement for Metal Plating

The pitchers to be plated are made the negative poles or cathodes, and suspended in the electrolyte, a solution of silver nitrate, for instance. The positive pole or anode is a strip of silver. The current goes through, decomposing the electrolyte, and depositing the metal on the pitcher. The thickness of the deposit depends upon the strength of the current and the time it runs.

EXPERIMENT 40

Purpose.

To show the effect of common acids on common metals.

Directions.

Use small pieces of the following metals: silver, copper, iron, brass, tin, nickel, zinc and aluminum.

a. Fill as many test-tubes as there are pieces of metal one-fourth full of hydrochloric acid (dilute). Drop a bit of each metal into each tube. Warm gently if necessary. Note what happens.

b. Fill another set of test-tubes one-fourth full of oxalic acid. Then drop a small piece of each metal into each tube. Warm gently. Note what happens.

Make a table showing the solubility of these metals in the acids.

Previous work with metals proved that in very many ways they may be attacked by various substances. They may be attacked by the oxygen of the air, forming variously colored oxides; other gases in the air may form compounds with the metals, especially carbon dioxide, which is always present, and hydrogen sulphide, which is present in certain localities; metals may be corroded by acids or other liquids, or they may be acted upon chemically by certain foods. All these agents combine to form the common ordinary tarnish with which most of us are familiar in one way or another. Metals also get dusty, and get covered with the greasy film variety of dirt. This general dusty greasiness, such as frequents the kitchen tea kettle, is easily removed by the application of the old stand-by, hot water and soap, followed by rubbing until dry.

Silver is the most common metal in household use. Whenever illuminating gas is used the sulphur compounds cause silver to tarnish very quickly. Some foods, because they contain sulphur, produce the same results. Sometimes silver wrapped in paper or cloth gets tarnished as easily as when in ordinary use. Nitrogenous compounds and sulphur compounds are often present in paper or cloth, and so of course the tarnish is produced in the same way.

Silver tarnish is removed either by mechanical rubbing off with certain kinds of silver polish or by some chemical action. Silver polishes have as their foundation material whiting (calcium carbonate) or diatomaceous earth (silicon dioxide). These solids are mixed either with alcohol or water or ammonia. These polishes of course rub off the tarnish, and their good use depends upon the fineness of the particles. The coarser the particles the more scratches are made and the more the silver is worn off. For silver ware which is plated this wearing off is much more serious, because the plating is only a thin coating of silver and the scratching or wearing off soon exposes the foundation metal.

There are several ways of removing silver tarnish involving chemical action.

EXPERIMENT 41

Purpose.

To clean silver.

Directions.

Either get or make some tarnished silver by bringing a dime or spoon in contact with exit tube of hydrogen sulphide generator.

a. Make a strong solution of washing soda or potash or borax. Put in the tarnished silver and boil for some time until clean. Let it stand until cool, then remove, rinse and dry with rubbing.

b. Take another tarnished piece and rub it with moist sodium chloride, then wash in ammonia water.

c. Get a small tin pan and two strips of zinc about an inch wide which will fit into the pan. Put into the pan hot water, to every quart of which has been added four teaspoons of salt and one of baking soda. Place the silver so it comes in contact with zinc strips and leave it until tarnish is removed. (Electrolysis takes place here.)

Questions. Which of these methods is quickest? Which will clean silver in large quantities easily? Explain action in *c*.

Write reactions for *a* and *b*.

The last method for cleaning is used commercially. The tin pan is replaced by rectangular pans made of zinc, one fitting inside the other. The inner pan is full of holes so that the silver may be placed in it; then put into the outer pan containing the solution of salt and soda. An electric current is set up between the silver, zinc and salts. This current then decomposes the tarnish of the silver.

EXPERIMENT 42

Purpose.

To study the composition of silver polishes.

Directions.

Get samples of the ordinary silver polishes.
Get tarnished silver as in preceding experiment.

a. Dissolve in water. Filter and test filtrate with litmus.

b. Test a portion of residue on filter with acid to find whether it is a carbonate.

c. Test another portion for silica.

d. Test unsoluble residues for fineness.

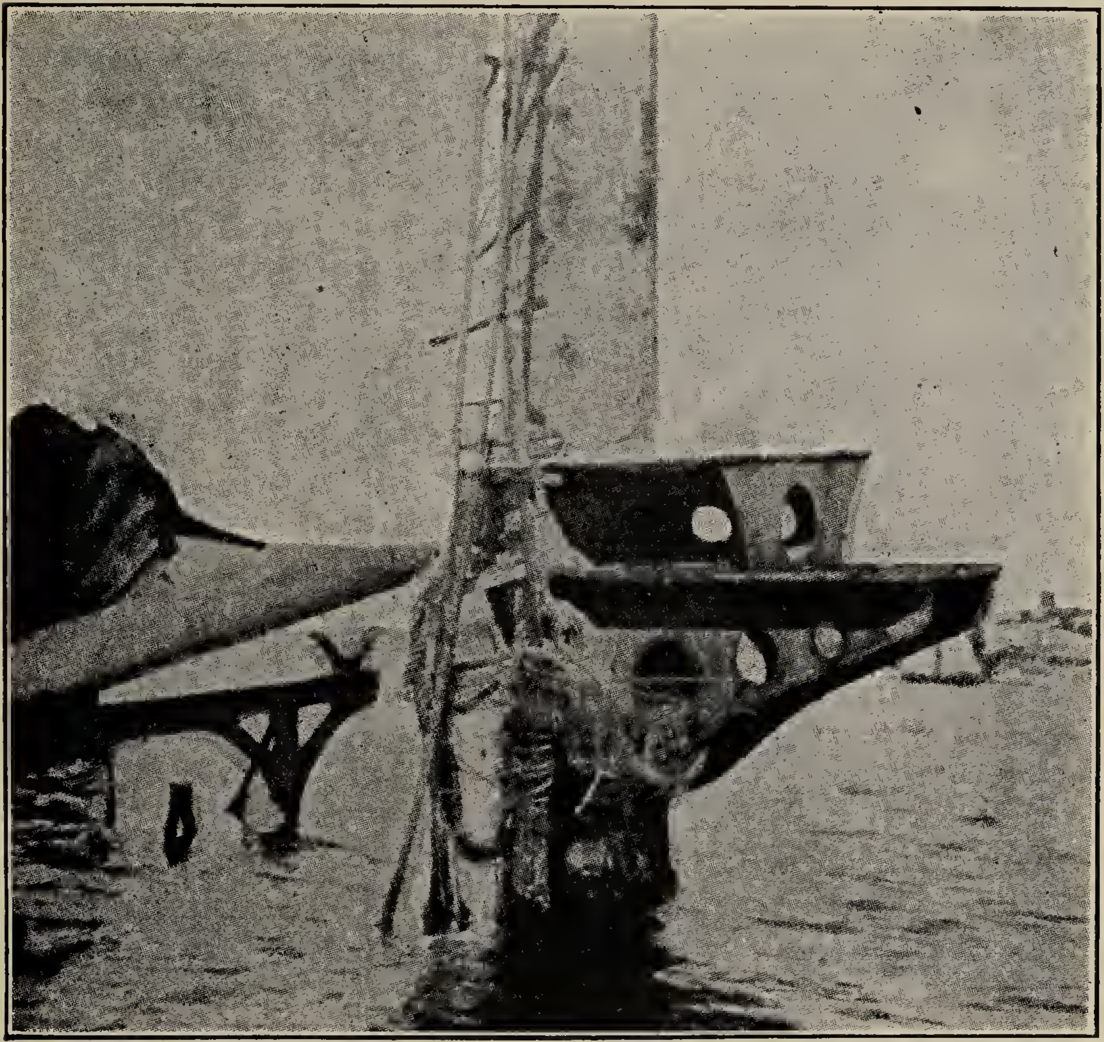
e. Try each polish on a bit of silver and compare them.

f. Make a paste of chalk or whiting and ammonia. Try it as a silver cleaner.

Questions. Are these polishes acid or alkaline? Gritty or smooth? Does the method in *f* work as well as the others?

Iron and steel rust very quickly in the presence of damp air. Iron rust is probably a combination of iron oxide, hydroxide and carbonate. It scales off in the form of a powder, thus constantly exposing new surfaces. This rust may be removed by the use of emery and oil or by rubbing with hydrochloride acid. In case the acid is used much care must be taken to remove all the acid. Washing with ammonia helps to remove all traces of the acid. Articles made of iron ought to be covered with a thin coating of some oily substance, especially if the article is not in constant use.

The tops of kitchen stoves, because of so much heat near them, offer good chances for oxidation, and a stove left untouched will finally become reddish color, the typical color of iron rust. Most stove polishes are made of graphite. Graphite burns very slowly and without odor and will when rubbed take a good polish. Spots



From "Municipal Chemistry"
Edited by Baskerville, published by McGraw-Hill Book Co.

Exposed Steel Platform U. S. Battleship *Maine*

A steel deck which was originally perhaps one-half an inch thick, appeared to be still intact, but on attempting to climb up on this piece of steel, it gave way under the ordinary pressure of the hand, showing that it had corroded entirely through.

made by the spilling of food cannot be so easily removed from the top of the stove polished with graphite as from one rubbed with a soft cloth and kerosene, but the rubbing with a soft cloth and kerosene does not produce so fine a polish.

Copper and its alloys brass and bronze are in common use in plumbing arrangements, and occasionally as cooking utensils, but the production of cheaper agate ware and recently of aluminum has relegated copper cooking utensils to the top shelf. Copper, as well as its alloys, is like iron attacked by the oxygen, water and carbon dioxide of the air, and so the tarnish is a mixture of all the compounds made. All of the compounds are soluble in acids, and so such common acids as lemon juice or vinegar, cream of tartar, buttermilk, may be used. Acids like oxalic and citric may also be used. Oxalic is poisonous and therefore not especially desirable to have in a house. Ammonia will also dissolve the oxide of copper, and this makes a good cleansing agent. All of these solvents must be carefully removed so as to prevent further action on the metal.

EXPERIMENT 43

Purpose.

To show effect of ammonia on copper oxide.

Directions.

Put a pinch of powdered copper (cupric) oxide in a test-tube and fill one-third full of ammonium hydroxide. Allow this to stand some time (perhaps two days or more).

Questions. Does the copper oxide go into solution? Is it rapid or slow?

These metals can of course be cleaned by the abrasive method. Rouge, Venetian red, whiting, putty powder and rottenstone (tripoli) may be used. These abrasives are softer than the copper or brass and so with the pressure of rubbing put a sort of film on the metal.

Nickel, because of the ease of electric plating, has come into much use as trimmings on stoves, coverings for pipes, and for some table utensils. Nickel does not oxidize easily under ordinary circumstances, but will become tarnished where gas is used. It may be kept clean by the use of hot soapy water or by some of the finer scouring materials. The same materials can be used for nickel that are suitable for tin. Aluminum is now used extensively for cooking utensils. One of its greatest advantages is that it is not attacked by ordinary substances and therefore does not tarnish easily. It sometimes darkens, especially with salt, but such spots are easily removed by fine powders. Alkalies tend to darken aluminum so that cleaning with some scouring powders is not to be recommended.

EXPERIMENT 44

Purpose.

To study action of various substances on aluminum.

Directions.

a. Into a small aluminum pan put some salt and water. Boil gently for three or four minutes.

(A vegetable containing salts may be used.)

Question. Are spots made on the metal?

b. Try to remove this spot by the use of a white cloth and Dutch Cleanser or some alkaline cleanser.

Questions. Does the spot come off? Note the color of the cloth.

Zinc and tin are both used as protective coverings for iron, giving the familiar galvanized and tin ware. Tin does not oxidize, zinc tarnishes slowly, but the oxide is white and therefore does not look so badly, even though the zinc loses its luster. The cleaning of these metals is the simplest method possible, soap and water.

Tin to serve its purpose must not be cleaned with a course abrasive, because the underlying iron oxidizes much more easily than untinned iron. When zinc wears off the rusting does not take place so easily because the zinc has formed some sort of a protective covering. For this reason galvanized iron is much to be preferred for out-of-doors ware.

Most of the metal polishes contain an abrasive of some kind held together with an alkali and some hydrocarbon. The presence of the alkali and the hydrocarbon help in the removal of the greasiness which is apt to get mixed with the tarnish.

Summary. Ordinary greasy dirt is removed from metals by hot soapy water. Tarnish is removed from iron best by kerosene; from silver by a simple fine polish like whiting and ammonia; from copper by ammonia, dilute acid, or fine abrasive; from brass by a fine abrasive; from aluminum by a fine abrasive, preferably one that is not too alkaline.

QUESTIONS

1. Why does alcohol clean eyeglasses?
2. What are the common abrasives for metals?
3. What is the chemical action involved in cleaning metals with acids?
4. Account for the green color on copper exposed to moist air.
5. Why does a spot of rust form on a tin pan?
6. What harm will a coarse polish do to solid silver? To plated silver?
7. How should silver be packed?
8. Why will kerosene remove iron rust?
9. Why does rubbing with kerosene keep copper and brass articles bright?
10. Why does household cutlery, if kept polished and dry, not corrode or tarnish so easily?

11. What is lacquer? Why is it used?

12. Why is kerosene better to clean windows in winter than soap and water or ammonia and water?

13. What is the chemical name of silver tarnish? Write reaction.

14. Why should galvanized iron not be used for cooking utensils? (See solubility of metals.)

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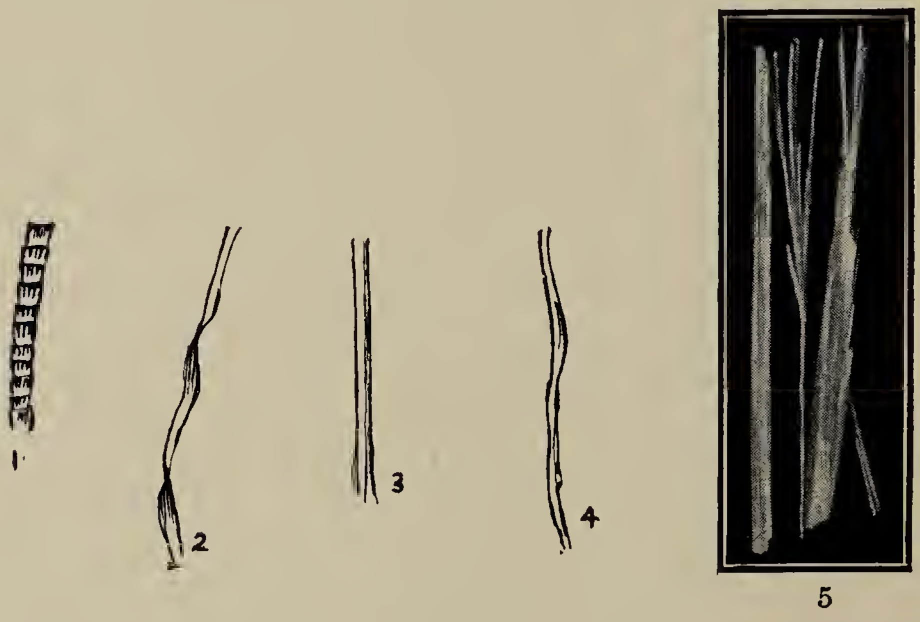
CHAPTER XV

FABRICS AND THEIR CLEANING

THE cleaning of fabrics is a much more complicated process than the cleaning of anything else. In the first place the varieties of soil are more numerous, and in the second place the cleanser will attack the fabric itself chemically as well as the soil. The problem then is first to know the nature of the dirt, then the kind of fabric. Both will have to be considered before the application of the reagent, for the reagent which will remove the soil might also destroy the fabric.



1. Wool fiber. 2. Cotton fiber. 3. Silk fiber. 4. Flax fiber. 5. Asbestos fiber.

The number of fabrics is at the present time very large, but the textiles which go into the weaving of these fabrics are very few in number. A textile is any material that can be woven into a fabric. It consists of fibers which must have the qualities of firmness, length, elasticity, strength, porosity, thinness, and uniform size in order to be made into thread which can be woven.

There are a great many fibrous substances in the world, but very few of them have the above mentioned qualities, or if these qualities are present the available amount of such fibers is limited and so for commercial purposes they are valueless. Fibers may be classified as follows:

Animal fibers such as wool, silk, mohair.
Vegetable fibers such as cotton, flax, hemp.
Mineral fibers such as asbestos, tinsel.
Artificial fibers such as spun glass, artificial silk.

EXPERIMENT 45

Purpose.

To study four kinds of fibers.

Directions.

Procure samples of raw wool or undyed yarn, raw cotton or unbleached cotton fabric, raw silk, linen thread.

a. Separate a bit of each substance into fine threads with a needle. Mount and examine with the compound microscope in the usual way. Compare the structure of the fibers.

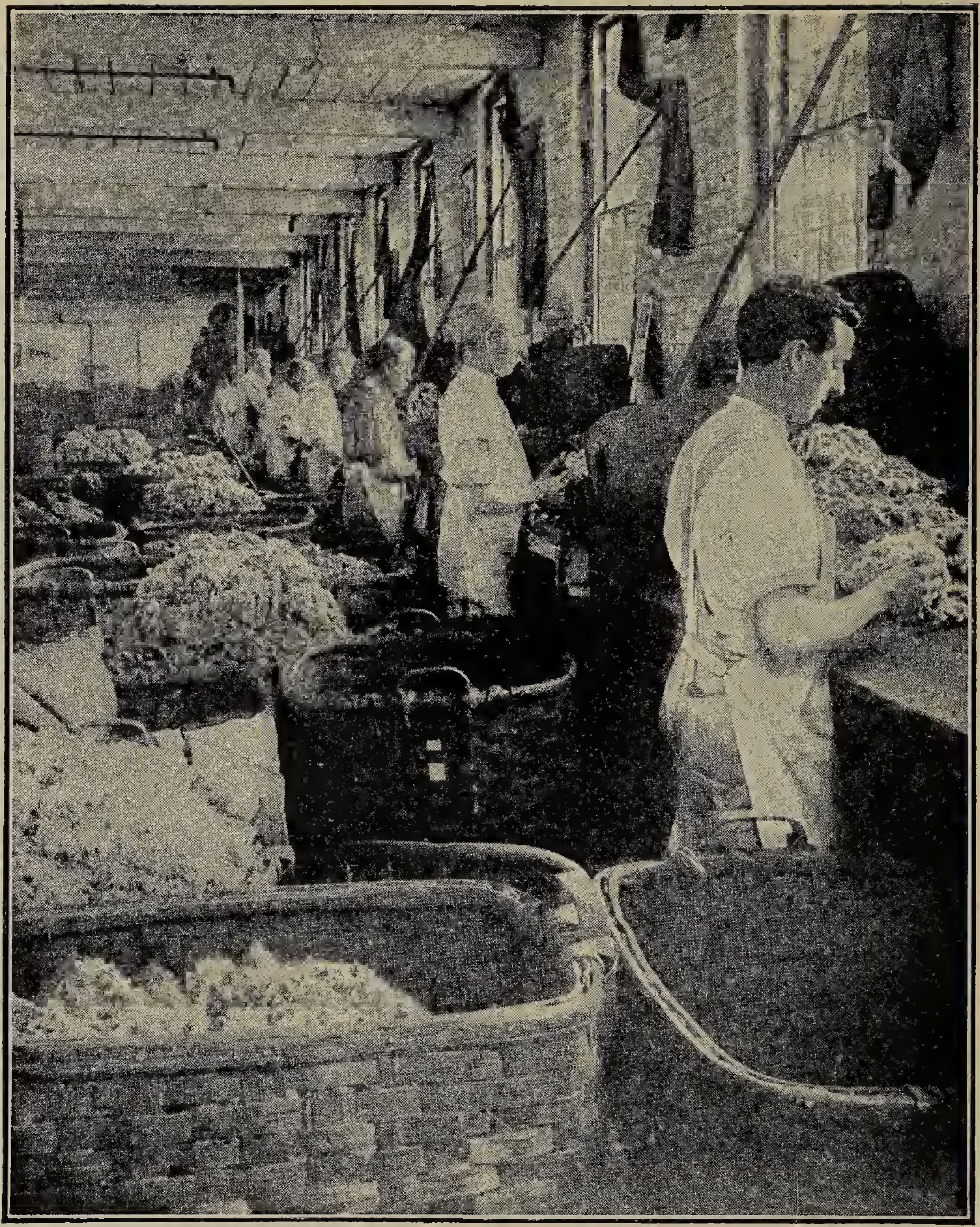
b. Pull out each kind of fiber to its full length, then release. Compare elasticity.

c. Hold by means of forceps some of each fiber in Bunsen burner flame. Note speed of burning and kind of ash left.

d. Break each fiber. Note tensile strength. Compare appearance of broken ends.

e. Try solubility of each fiber in concentrated caustic soda (hot and cold); hydrochloric acid (dilute and concentrated) (hot and cold); nitric acid (dilute and concentrated) (hot and cold); sulphuric acid (dilute and concentrated) (hot and cold); ammonia water; borax and water; washing soda.

f. Lay a piece of all cotton fabric and all linen fabric flat on top of some water in an evaporating dish. Compare action in the two cases.



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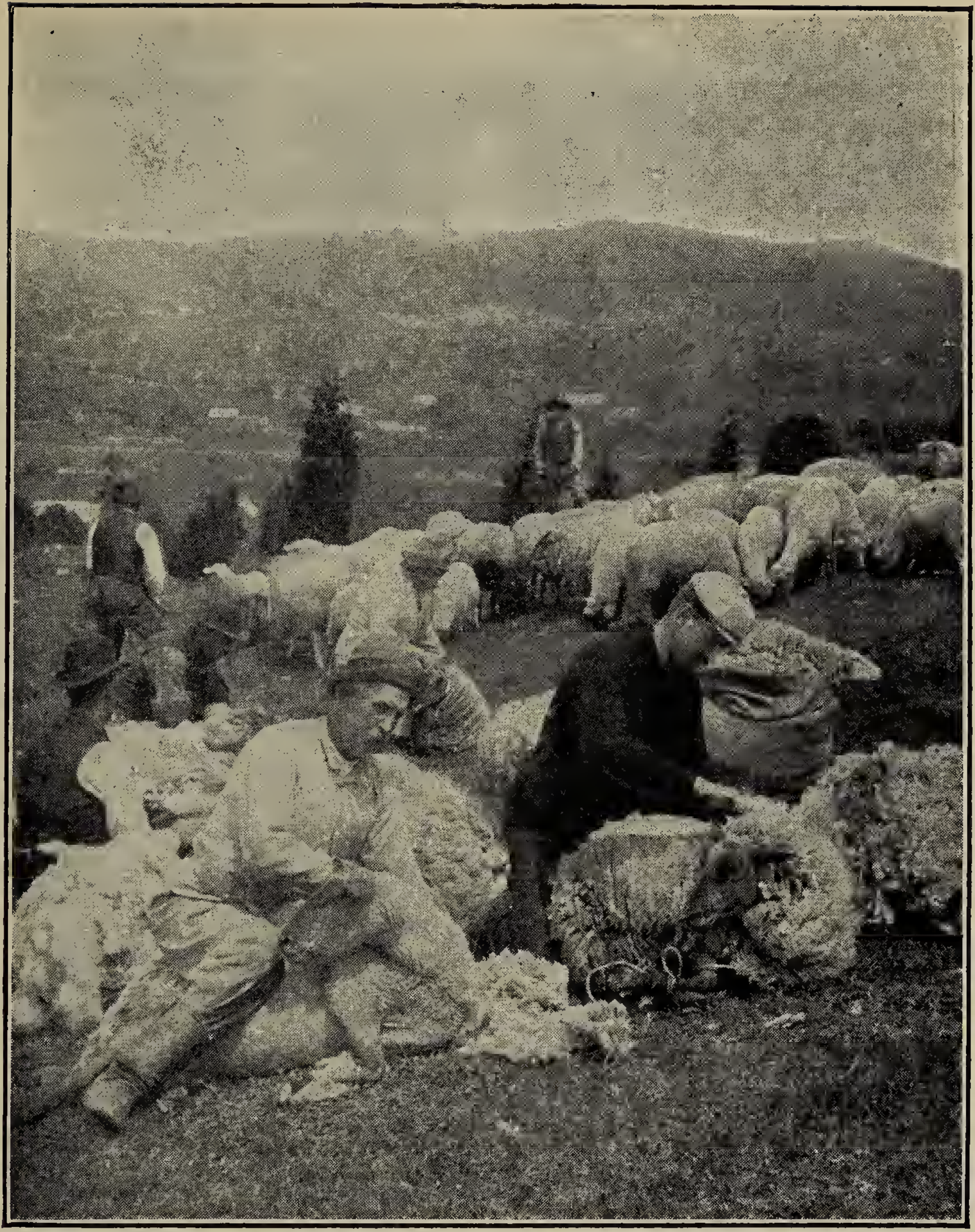
Sorting Raw Wool into Three Grades in an American Woolen Mill

g. Stain a piece of cotton fabric and linen fabric with fuchsine. Note the color. Then add ammonia water to each.

Wool is the fleecy covering of the sheep, goat or alpaca. Different kinds of sheep produce different kinds of fleece, which gives rise to the varieties of wool. Different parts of the same animal body produce different qualities; for



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Shearing Sheep on an Eastern Sheep Farm

instance, the covering from the head and neck of the animal gives the best grade of wool. The wool fiber varies from one to eight inches in length, and under the microscope appears to be covered with a number of tiny scales which project outward, giving it a roughened appearance. When the fibers are placed together to make thread

Ⓒ Underwood & Underwood, N. Y.

Feeding Silk Worms their Breakfast of Mulberry Leaves, Mt. Lebanon, Syria

these scales interlock. This interlocking gives wool its chief characteristic, a felting or shrinking power. The more closely these scales interlock the greater the felting power, and the closer the structure of the thread becomes. Wool chemically is a nitrogenous compound containing sulphur.

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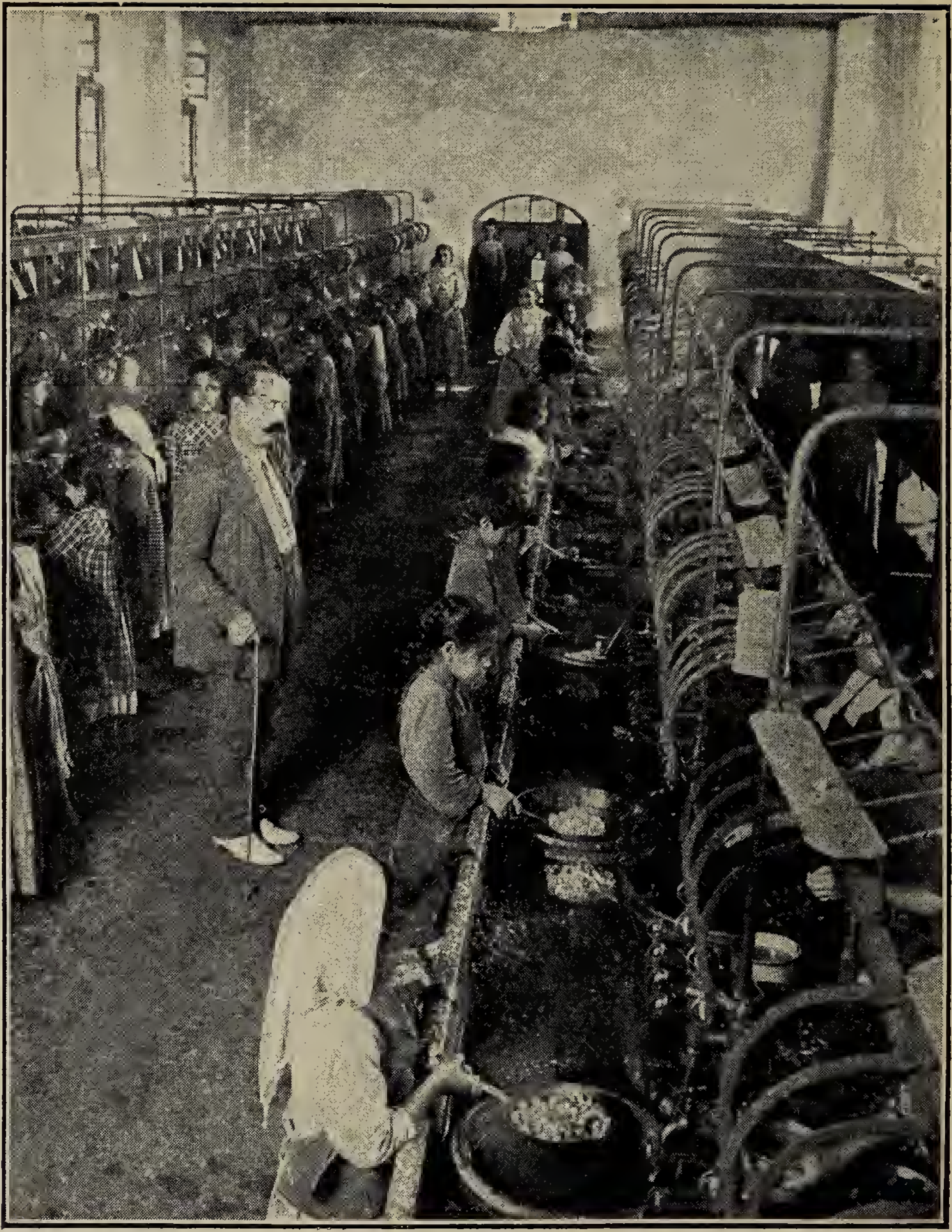
Up-Country Independent Silk Growers coming with Bags of Cocoons for Sale. Syria

The man in the center is an export agent who buys for the French trade. The fibers holding the cocoons together can be seen easily.

Silk is another animal fiber of great commercial importance. It is a material secreted by an insect in the caterpillar stage to protect itself during the development into the adult moth. The cocoon thus made has to be kept intact so that the fibers may not be broken. The



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Boiling Cocoons to Loosen Fiber Ends in Syria's Largest Silk-Reeling Plant, Mt. Lebanon

Note the children at work.

pupa, as the insect is called during this protected stage, is killed by steaming, the fibers are softened and separated, and then woven into threads. The fiber under the microscope is smooth and cylindrical. It is a nitrogenous substance, like wool, but contains no sulphur. It is a much longer fiber than wool, being from 1000 to 4000 feet long. It differs from wool in that as it is made by the animal it is already a thread.

There are some seed coverings of a fibrous nature which possess a luster almost like that of silk. Some of these have been used as silk, but the fibers are not so long or strong and the result is not specially satisfactory. These seed coverings are like cotton made of cellulose, and because of this attempts are made to produce a silk from waste paper, getting by chemical action pure cellulose from the waste paper or wool. Cellulose can be made into collodion, popularly known as "new skin," and in this form is expelled through small tubes. After drying and further treatment it can be spun like the animal fiber silk. It has a luster quite like natural silk but is not so strong.

EXPERIMENT 46

Purpose.

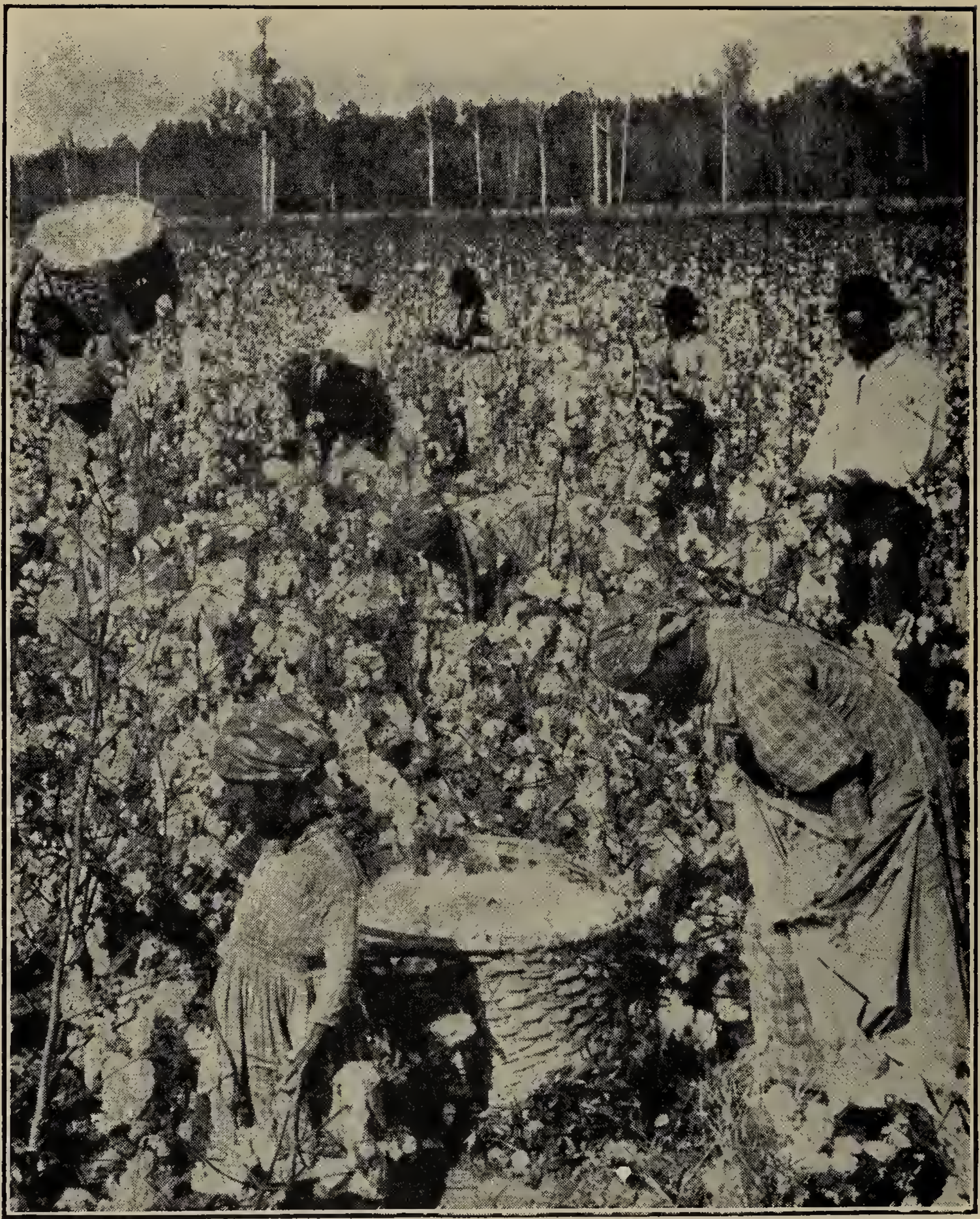
To distinguish chemically between artificial and natural silk.

Directions.

Get various samples of silk fabrics, especially cheap silks (preferably undyed silks). A pure silk boiled in a four per cent solution of potassium hydroxide produces a colorless solution. An artificial silk boiled in the same kind of solution produces a yellow solution. Try the various samples procured to see which is animal, which vegetable fiber.

Questions. Why is it better to have the samples undyed for this experiment? How would you test a dyed sample?

Cotton is the most important vegetable fiber used for weaving fabrics. It is a soft downy substance which protects the seed while still in the fruit or boll. The



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Cotton Picking

cotton fiber is shorter than the silk or wool fiber and is characterized by a peculiar twist in it. This twist helps interlock the various fibers as they are spun into thread, in much the same way that the scales of the wool fiber help.

Cotton grows best in warm climates, and there are therefore many varieties according to the nature of soil in which it is grown and the care which the crop receives. Cotton chemically belongs to the class of carbohydrates known as celluloses.

A cotton fabric may be so treated as to get a glossy appearance like linen, a more expensive fiber. The threads of a cotton fabric are more nearly uniform in size in both warp and woof, while the linen threads are uneven. This sometimes helps in detecting one or the other without chemical tests.

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Flax from Stacks Soaking in River Lys, Preparatory to Spinning. Courtai, Belgium

Linen, like cotton, is a vegetable fiber and comes from the bast tissue of the stem of the flax plant. It is, under the microscope, a long cylindrical fiber with no twist. It averages a length of twenty inches and has a greater tenacity than cotton. It has a smooth texture and luster which when the fiber is woven is increased by laundering. Like cotton it is chemically cellulose.

The ordinary soil of underclothing comes from the waste products of the body, chiefly perspiration and oily substances. Hot water, soap and rubbing remove this without much difficulty. Most authorities agree that the best laundering results are obtained when little water and much soap are used, followed by careful rinsing in clear water. Even with ordinarily soft water, good soap, and attention, white clothing will become yellowish. Sometimes this is due to substances used in bleaching the fabric, sometimes to the formation of substances by the action of soap on the dirt. This is usually eliminated by rinsing in bluing water. Bluing, as its name implies, is a blue coloring matter, not a dye. Since blue and yellow are complimentary colors the yellow is covered with the blue color, and white is the result.

The substances which make bluing are ultramarine, containing sodium, aluminum, silicon, sulphur and oxygen; indigo, coal tar products, and Prussian blue, an iron compound (formerly made from plants, now from coal tar). Ultramarine and indigo are insoluble in water and so have to be deposited in fine particles over the laundered article. The coal tar products are good stable substances, are soluble and therefore can be spread more satisfactorily over the garment. Prussian blue is soluble and cheap. It is, however, easily attacked by alkalis, forming iron hydroxide. This hydroxide when subjected to the heat of the ironing process is apt to make a brownish rust spot. If, however, the solution used is not too strong, and if it is spread evenly through the rinse water, rust spots are usually avoided.

There are however other sorts of soil that give much more trouble than the preceding; spots that get on to clothing by carelessness or accidents, and while they may not be very large in extent are annoying to a degree that is inversely proportional to their size. This variety of soil includes many kinds of substances both organic and inorganic. The organic soils and stains include the one mentioned above, also blood, egg, milk, butter or cream, fats, stains from fruits and vegetables, and grass, molds and mildews. The inorganic soils and stains are produced by medicines, inks, paints, wagon grease and machine oils. With many of these, especially if they are oily or greasy, dust is mixed, which tends to make the spot more evident, though it does not especially complicate its removal.

The substance used for the removal of spots and stains from fabrics acts as a solvent, an absorbent, a detergent or a bleach. The solvent action may or may not be chemical, the absorbent action is purely physical, the detergent action may or may not be chemical, but the bleaching is always chemical and oftentimes a very complicated action. Any one of these methods or combinations of them may be used according to the need.

The solvents as such have been discussed in a previous chapter, and also their action on fibers. It will be well to review the work at this stage. Water is the most common one. However, from ordinary experience one knows cold water alone removes very little dirt. Mildew of recent formation may be removed by plain cold water, but if it is an old stain water will have no effect. Grass stain, milk and cream may also be removed by cold water. Warming the water increases its power, and at the boiling point water alone may do a great deal. Fresh blood stains, glue and syrupy stains may be removed by washing in warm water. Boiling water will remove bluing stains, some fruit and coffee stains. Boiling water has its best effects when applied to the fresh stain. None of these spots can be removed by this method from fabrics which are hurt by water. Spots on fabrics such as silks can be removed by the water method if they are not too large, and if the garment or material is shaken vigorously in steam. An easy way to do this is to shake the material in the steam issuing from an ordinary teakettle until it is thoroughly moist. Then shake it thoroughly dry.

A detergent like soap, ammonia or borax added to water increases its power of removing spots. Blood stains may be removed easily by warm water and naphtha soap, or by warm water and ammonia; chocolate is removed by cold water and borax; milk and cream spots and meat juice spots, after being washed with cold water, ought to be washed with warm water and soap. Many oils like kerosene are taken out by warm water and soap; in fact that is the method for removing the ordinary soil from clothes in the weekly wash. A detergent is any substance that cleans. The most common ones have been mentioned in connection with water. The others are really solvents like benzene and naphtha and have been discussed with the solvents.

The absorbents act on spots merely in a physical way, absorbing the foreign material from a fabric while the material is still fresh. The most common ones are chalk, magnesia, Fuller's earth, starch and meal, and porous papers like blotters. The best method is to apply the absorbent on both sides and to remove it as fast as it absorbs.

The best and most harmless bleaches are oxygen and sunshine. Our grandmothers depended very much on this to whiten their clothes and fabrics. Especially was this true when most fabrics were made from the natural fiber. This process is however a long one and, in the modern rushed and hurried living, most people are unwilling to wait. Then too the increased number of substances which can make spots adds greatly to this need of a bleach that will work more quickly than fresh air and sunshine. The most common bleaching substances are sulphur dioxide, borax, ammonia, bleaching powder (chloride of lime), oxalic acid, sodium hypochlorite and hydrogen peroxide. Of these the sulphur dioxide is most difficult to use because it is a gas which has an unpleasant effect on the nose and throat. Hydrogen peroxide is excellent but is too expensive. Borax and ammonia are mild bleaches, especially if used in the final rinse water.

Bleaching powder depends for its action on the chlorine set free, the hydrogen peroxide for its oxygen. Chlorine acts on water, setting free the oxygen, and so ultimately these bleaches involve the action of oxygen. This is also true for sodium hypochlorite. Their preference over fresh air and sunshine is that the oxygen is set free in atomic condition and is therefore more active. Javelle water is a preparation containing soda and chloride of lime, so its action is similar to bleaching powder, which is chemically an oxidizing process. Oxalic acid has a feeble reducing action, and so reduces the complicated substances containing color to simpler substances having no color.

EXPERIMENT 47

Purpose.

To study the effect of various bleaches on textiles.

Directions.

Procure samples of pure silk, wool, cotton and linen fabrics, *with no color*. Try each in a solution of the various bleaches mentioned in text. Allow one set to stand over night, then rinse carefully in cold water. In another set boil the fabric with the bleaching solution. Rinse carefully in cold water. Compare the samples that have been treated with those that have been untouched.

NOTE. Directions for making Javelle water: 1 lb. washing soda; 1 qt. boiling water; ½ lb. chloride of lime; 2 qts. cold water.

Question. What are your opinions as to the effect of the bleach on the fabric?

EXPERIMENT 48

Purpose.

To try various bleaching agents on stains.

Directions.

a. Procure samples of various fabrics and put on them spots of various kinds; fruits, coffee, ink, blood stain, grass, etc. (according to class needs). Apply the various bleaching agents used in previous experiment.

Questions. Tabulate results. How does the actual work compare with textbook statements?

A bleach is the last substance to be tried in the removal of soil. With a white material it is not so much of a chance, but with a colored fabric the chances are that the color goes with the spot; and sometimes the whole fabric or garment must be bleached. All bleaches should be applied in small quantities and should be followed by thorough rinsing. There are a few simple rules to follow when trying to get rid of spots or stains. These may briefly be stated as follows:

1. Know the material.
2. Know the nature of stain if possible.
3. Use dilute reagents (except organic solvents) and in small quantities.
4. Rinse thoroughly either in warm water or in fresh amount of the organic solvent.
5. Use the simplest substances that will attack the

spot, first washing between applications of different reagents.

When using a small quantity of reagent for the removal of a spot or stain it is best to hold the fabric tightly over a bowl or place the spot on white blotting paper. This helps prevent the ring so commonly made by naphtha or gasoline. In cleaning a whole garment it is better to dip repeatedly in the cleanser without rubbing.

EXPERIMENT 49

Purpose.

To remove a stain from a fabric.

Directions.

Various fabrics with spots of unknown origin are to be used, trying to remove the spot as satisfactorily as possible.

QUESTIONS

1. How could you determine whether a fabric is all wool or a mixture of wool and cotton?
2. How can you tell whether a fabric is all linen or a mixture of linen and cotton?
3. Give a chemical test for wool, for silk, for cotton, for linen.
4. What property of wool makes it not so satisfactory for babies' underclothes as a mixture of silk and wool?
5. Give two reasons why a thin glaze should be on china.
6. Why does a housewife need to know chemical actions of textile fibers?
7. Why is the heating of wool by hot water like heating white of egg?
8. What makes wool shrink?
9. What is "mercerizing"?
10. Is Javelle water suitable to remove a stain from colored cloth? Why?
11. Why will a washing compound usually change the color of a cloth?
12. Why are clothes put to soak in cold water?

13. Why are blued clothes often streaky?

14. Why do blued clothes sometimes turn yellow?

15. Why do fruits stain colored clothes?

16. Why is a silk and wool mixture easier to clean than a silk and cotton one?

17. Why is it not wise to first use hot water on a milk spot?

18. Why will a garment bleach better in sunshine if it is wet than if it is dry?

19. Why will ammonia remove grease spots from fabrics? Why is it a good substance to use?

20. Why is it not wise to rub a blood stain on a blanket with soap?

21. Find out by experimenting the nature of ink eradicators.

22. How does printer's ink differ from other kinds of ink? How could it be removed?

23. Why is it necessary to do so much washing and rinsing in the " removal of spots " experiment?

24. Why is Javelle water kept in a dark place?

25. Industrially what fibers are bleached? What is the reason for this?

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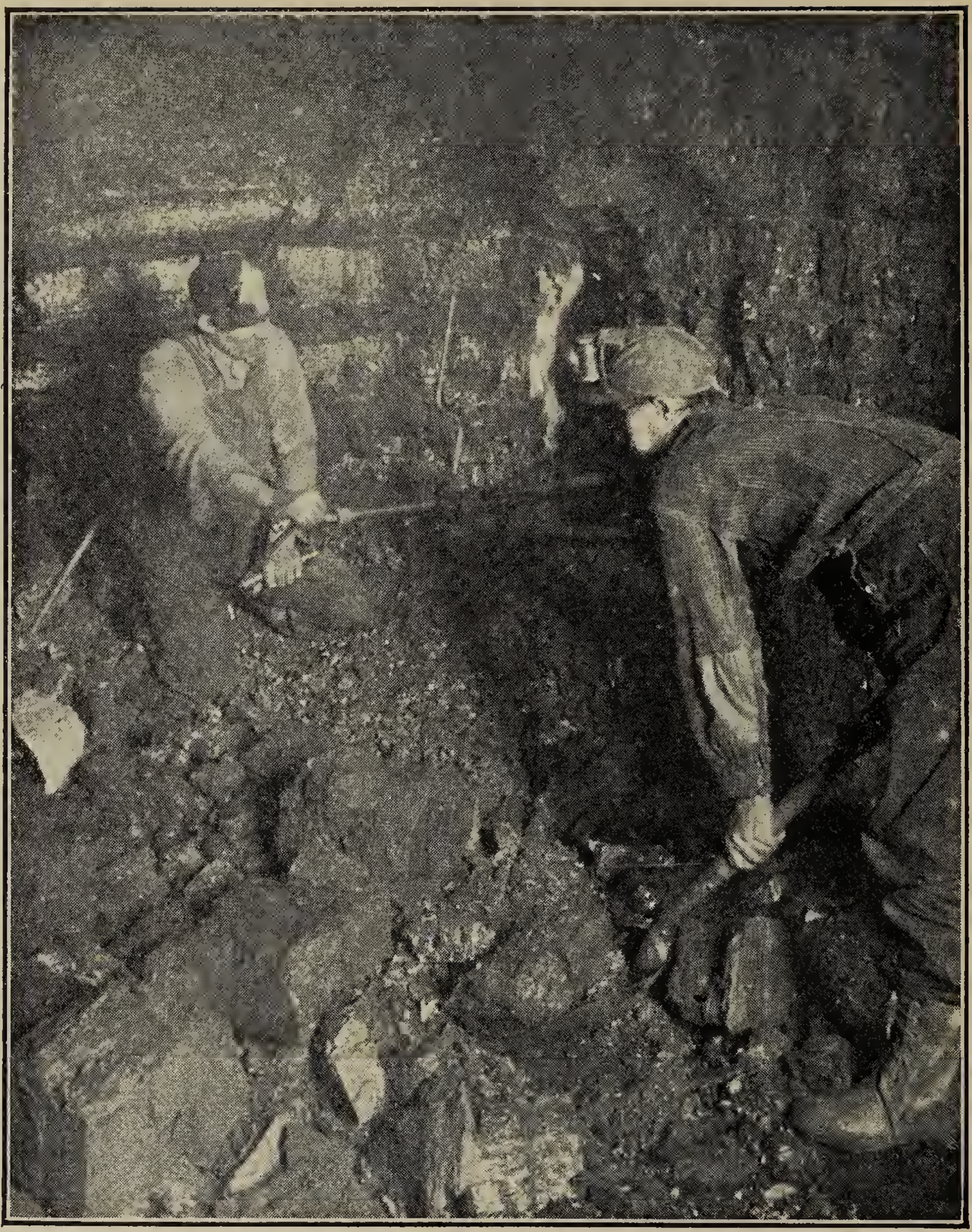
SUPPLEMENT

FUELS

A FUEL is any substance which when burning produces a large quantity of heat, little smoke, little ash, and which does not burn too rapidly. The most common fuels may be solids like wood or coal, liquids like kerosene or gasoline, or gaseous like natural gas or coal gas. In any case they are substances containing a great deal of carbon combined with oxygen and hydrogen, as in the case of woods; or combined with hydrogen only in the gasoline and coal gas. Their use depends upon the fact that when heated to a certain point called the " kindling temperature " they are decomposed into simpler substances which are then oxidized by the air. This oxidation sets free large amounts of heat.

Wood may be either hard or soft according to the structure of the tree from which it is taken. Most of the evergreen trees furnish soft wood, while trees like oak, chestnut, hickory and birch furnish the hard woods. The soft woods take fire easily and burn quickly, therefore liberating heat quickly. The harder woods do not burn so easily, but the heat production spreads itself over a longer period of time.

The two alcohols, methyl (wood) CH_3OH and ethyl (grain) C_2H_5OH, are used as fuels in a small way, such as in chafing dishes. The expense of production prevents a wider use. Wood alcohol, along with acetic acid and charcoal, is produced by the destructive distillation of wood. It is not so good a fuel as ethyl alcohol. Because of its poisonous qualities it has become necessary to " denature " it by putting in a substance which gives it a disagreeable odor. Ethyl alcohol is made by the fermentation of starches or sugars by certain enzymes. Since grains contain much starch they may be used as a starting



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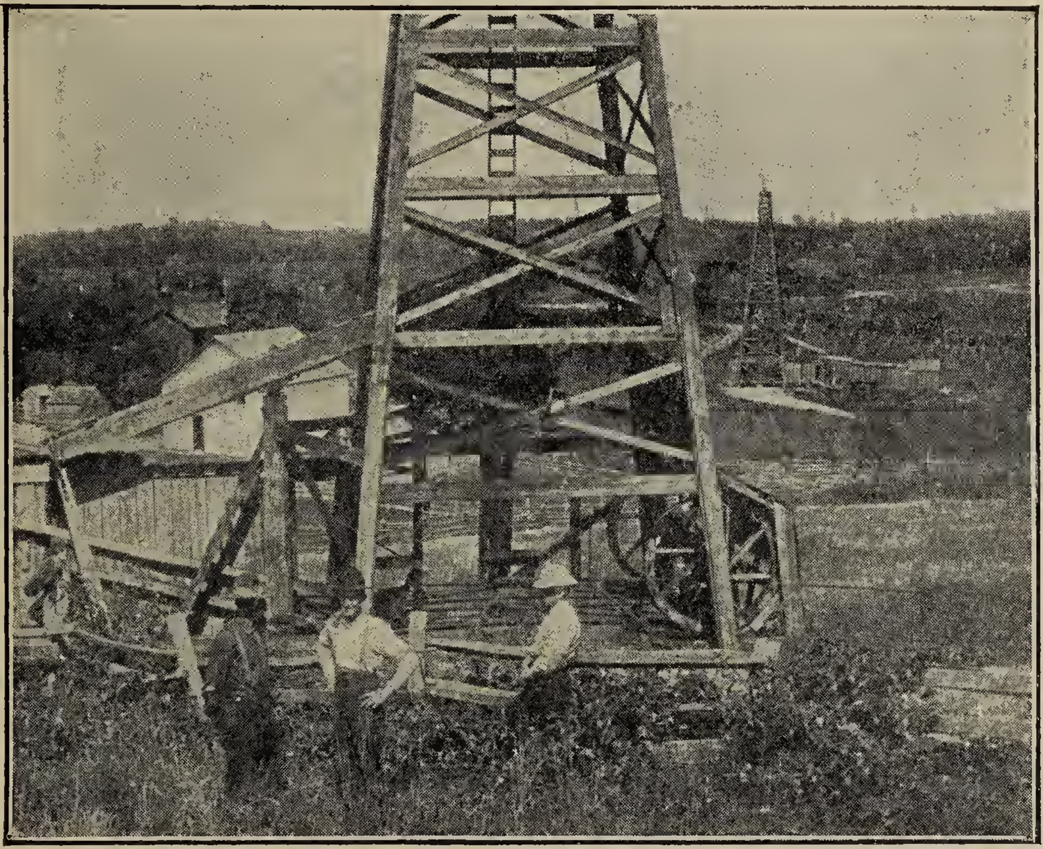
Coal Strata

Miner drilling for a blast in an anthracite mine, Scranton, Pa.

point, though the sugars of fruit do quite as well. All fermented and distilled liquors contain alcohol in varying proportions. Ethyl alcohol is the one commonly used in the household for medicinal purposes.

Coal like wood is of vegetable origin, but the origin was far back in geologic times. In the carboniferous age plants were much larger than those of the present, and in the great earth movements were buried many feet under the surface. Away from air and subjected to more or less heat a slow destructive distillation set in which resulted in the decomposing of the plant body so as to leave only the carbon in large amounts. The various stages of decomposition give the various grades of coal, lignite containing the least carbon, bituminous or soft next, and anthracite or hard containing the largest amount of carbon. The soft coals are used for industrial purposes while the hard coals are for domestic use. These vary from 50 per cent to 90 per cent carbon.



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Petroleum. Source of the World's Most Gigantic Fortunes. Pumping Wells in the Oil Country

Petroleum is a yellow-brown viscous liquid found in the earth in various places. It consists almost entirely of a mixture of hydrocarbons. When petroleum is subjected to a destructive distillation many important compounds are set free. Some of these are the liquid fuels. When liquid petroleum is heated the first distillate, whose boiling point is 40° to 70° F., is petroleum, ether and naphthas; the second, with boiling point 70° to 90° F., is gasoline; the third, whose boiling point is 80° to 120° F., is light petroleum used for solvents; the next distillate collected, with a boiling point of 150° to 300° F., is kerosene. The products collected above 300° F. are semi-solid, the most important substance made from American petroleum being vaseline. Paraffin is a solid substance produced from the Russian petroleum after the oils have been set free.

The naphthas, gasolines and kerosenes are all mixtures, and to produce a good grade have to be redistilled to a more definite specific gravity and flash test. Especially is this true for kerosene, whose flash point is 65.5° C. (150° F.) for the best kerosene. The flash point is that temperature at which the oil gives off enough vapor to maintain a continuous flame when ignited.

When soft coal is heated away from the air or destructively distilled, several important products are formed. One of these is ammoniacal liquor, which is the source of all the ammonia water. Another is coke, which is the residue after all other products have been collected. This coke is a grayish, light porous solid which is again used as a fuel. It burns, of course, much quicker than coal, but when pressed into briquettes makes a very good fuel. Still another is illuminating gas, which finds wide use both as an illuminant and as a fuel. It is largely a mixture of marsh gas and carbon monoxide. The presence of the latter makes the gas dangerous to breathe. Much of the gas of the present day is made not from soft coal but from passing steam through anthracite coal. Such gas is known as water gas and contains largely carbon monoxide and hydrogen. In some localities gas occurs naturally, issuing from the earth when borings are made. This gas is probably the result of more complete distillation of organic material which has gone on during past ages. A great deal of the natural gas in this country has been used, however, and its place has had to be filled by artificial substances.

EXPERIMENT 50

Purpose.

A study of the destructive distillation of soft coal.

Directions.

A clay pipe, soft coal, glass and rubber tubing, a U-tube, plaster of Paris. Put pieces of the coal into the bowl of the pipe and cover well with plaster of Paris. Clamp the pipe to a ring stand and heat. Heat the bowl of the pipe until dense fumes are given off.

a. Apply a lighted match at end of pipe stem. The gases should take fire.

b. Put out fire at end of pipe stem. Connect stem with U-tube half full of water by means of stoppers and glass tubes. Heat, and let gases pass through water in U-tube. Note what has formed in U-tube. Note odor.

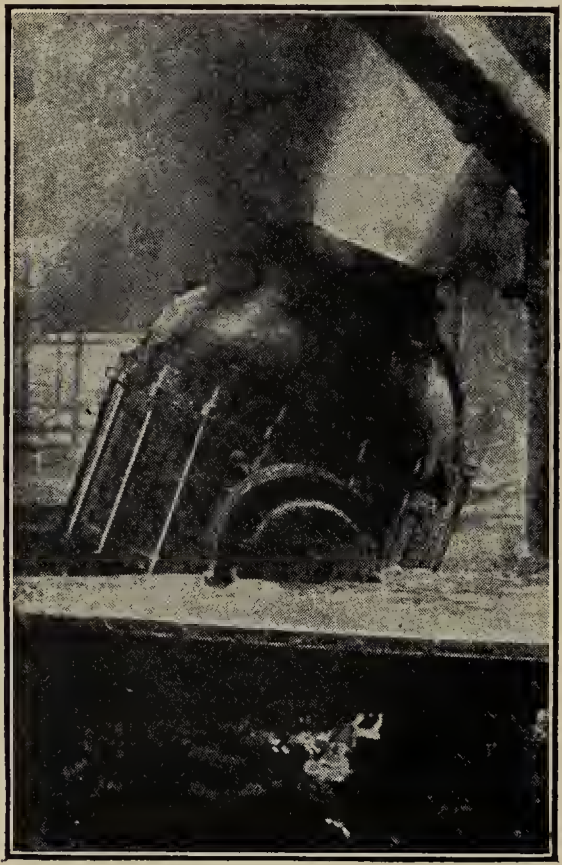
c. After heating has gone on for some time, disconnect. To the water in the U-tube add a little lime and heat. An odor of ammonia should be noticeable.

d. When the pipe has cooled, break it open. Compare appearance of contents with that of the original coal.

Question. What four varieties of substances are obtained by the decomposition of soft coal?

IRON

From the wide and varied uses to which iron may be put one would imagine it to be a very common element. Only about five per cent of the earth's crust is iron, so that all the iron existing as we know iron has to be made from compounds of iron. It does not exist free. The most important ores of iron are the sulphide (iron pyrite) FeS_2, the oxides, magnetite Fe_3O_4 and hematite Fe_2O_3 and the carbonate of iron produced from the oxides largely. The crude iron is made in blast furnaces by heating the oxides in a current or blast of air together with coke and a flux of some material which will combine with the impurities in the iron. This process of heating reduces the oxide to iron containing some impurities, and at the same time forming slag which is removed from time to time. The iron in a molten condition is run into molds forming bars of crude iron called " pig iron." Iron at this stage is the most impure form. This cast iron which expands upon cooling, contrary to the usual behavior of solids, finds wide use in the making of stoves, pipes, radiators and ranges. Cast iron is brittle, and cannot be welded or forged, and cannot be used for any object which may be subjected to shock or strain.


Bessemer Converter

Wrought iron is made from pig iron by burning out the carbon and other impurities as far as possible in reverberating furnaces. Pig iron may contain six per cent of these elements, but they are less than one per cent when wrought iron is produced.

Wrought iron is the purest commercial form of the metal. The removal of the impurities changes the properties of the metal entirely. It is much tougher than cast iron, can be forged and welded, and therefore is much used by blacksmiths and machinists.

Steel is a form of iron which stands between cast and wrought iron as to the amount of impurities it contains. For that reason iron in the form known commercially as steel, has different characteristics and therefore different uses. Steel used for watch springs and cutlery needs different qualities from those of a steel for girders, for instance. By varying the proportion of impurities between that of cast iron on the one hand and wrought iron on the other the desired results are produced. Large quantities of cheap steel are produced from pig iron by the Bessemer process. Another process giving better steel is known as the open hearth process and is fast coming into use.

Steel is hard and brittle and has the ability to stand tempering, which makes it a most valuable form of iron. It can stand much strain and stress, though the degree of all these properties depends upon the amount and nature of the impurity present.

ALLOYS

Just as iron containing varying proportions of the elements carbon, phosphorous, nickel, manganese, etc., changes its characteristics from cast iron to wrought, so the combination of different metals forms substances which vary from the characteristics of the original metal in such a way as to make most useful substances. Such a combination is called an alloy. The nature of such a combination is not perfectly understood. Some authorities believe it to be a chemical union, and some consider it a solution of one solid in another. Whatever their nature, the alloys form a very important class of substances. By changing the proportion of the constituent parts different degrees of hardness or brittleness may be obtained; or a different melting point is produced; or a better conductivity for heat and electricity may result.

Below is given a list of the most important alloys and their composition.

Brass:

60% copper, 40% zinc (sometimes 2% lead).

Bronze:

74.7% copper, 25.30% tin (sometimes zinc or aluminum).

Coins:

10% copper, 90% gold or silver.
75% copper, 25% nickel.
95% copper, 2.5% tin; 2.5% zinc.

Fusible metal:

28% copper, 72% silver.
50% bismuth, 27% lead, 13% antimony, 10% cadmium.

Solders:

Tin and lead (plumbers' solder).
Copper, zinc, tin.
Silver, copper, zinc.
Gold, copper, silver.

Type metal:

70 to 80% lead, 20 to 25% antimony, tin and copper.

White or Babbit metal (machinery bearings):

75.90% copper, tin, antimony.

EXPERIMENT 51

Purpose.

To make a fusible alloy.

Directions.

Weigh out 5 grms. of lead, 5 grms. tin and 10 grms. bismuth. Mix the lead in small pieces with the tin. Put the mixture into a porcelain crucible, support it on a clay triangle, and heat with Bunsen burner. When the mixture has melted add the bismuth. Stir with an iron wire as it heats until the mixture is entirely liquid and well mixed.

Pour the contents of the crucible while still molten into a dish of cold water to harden it. Heat a small piece in an evaporating dish of water and bring to boiling point.

Questions. Compare the alloy with original substances as to color, hardness and brittleness, melting point.

INKS

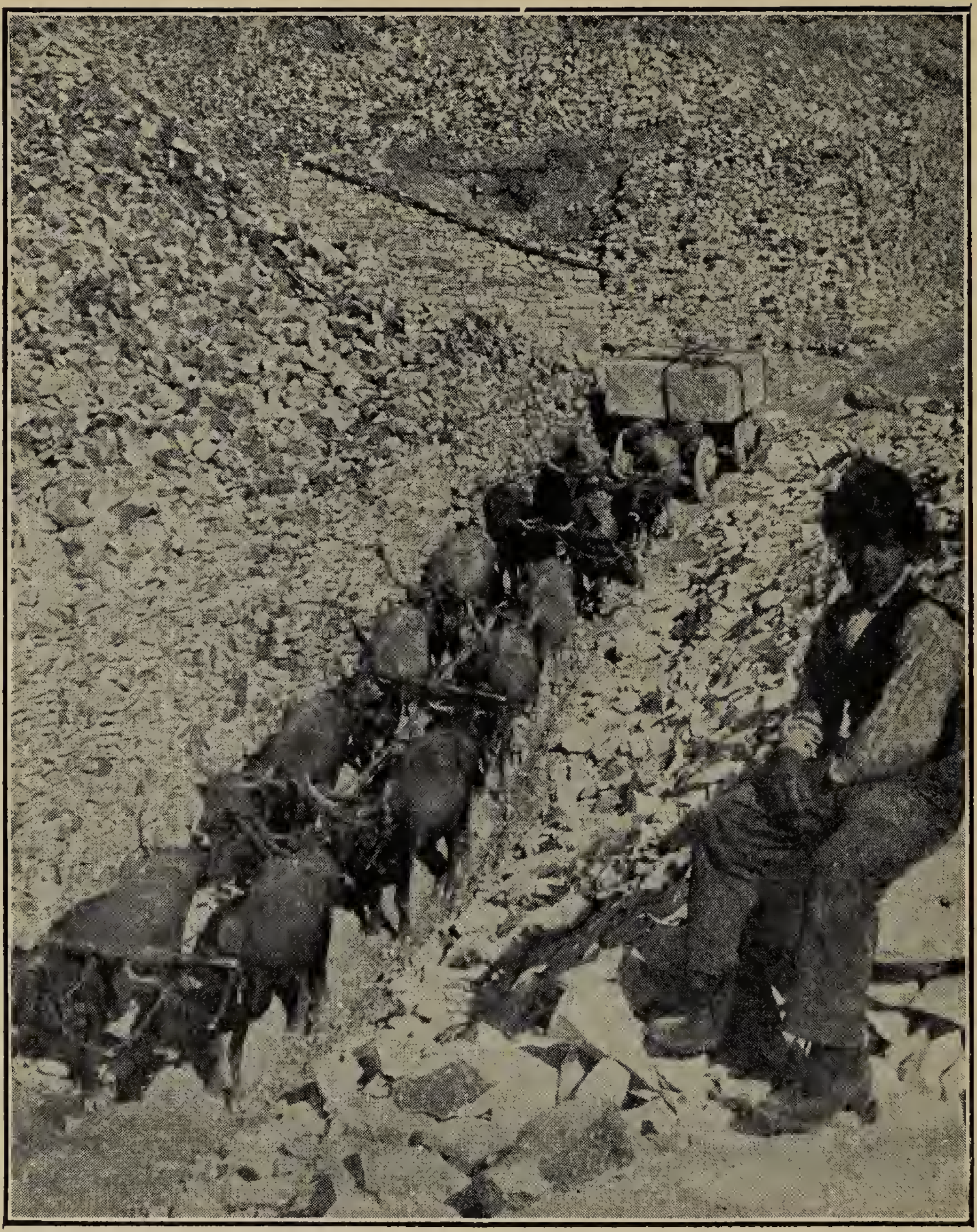
Inks may be used for writing, for copying purposes, for printing purposes or for making an indelible impression. They may be of any color, though the most commonly used inks are black. Because of their various uses they are made from different substances which produce the desired results. For writing purposes an ink should keep its color, flow readily, penetrate the fiber of the paper but not go through, must not corrode pens easily, must dry rapidly, and must not become moldy or thick. Writing inks are made in two ways, either from tannin and ferrous sulphate or by a water solution of some aniline (coal tar) dye.

The tannins are chemical substances which occur in the barks of various plants, such as oaks, chestnut and sumach. A tannin is a light brown, powdery soluble substance very astringent in its taste, having the power to turn animal skin into leather, and known chemically as gallotannic acid. The most common way of getting this is from the excrescences often seen on oak trees and known as oak galls. These galls are produced by the sting of a gallfly at egg-laying time. This sting naturally injures the tree so that the tree exudes a protective covering which forms the gall. These galls are crushed and the essential substance, gallotannic acid, extracted by ether.

Ordinary ink is made by the addition of ferrous sulphate or chloride to the gallotannic acid, and this acid hydrolyzed by acid into gallic acid. This gives no color reaction immediately, but when exposed to the air the ferrous salt of the gallic acid is oxidized to the ferric salt, which is a dark color. In order to produce a dark ink upon its immediate use some aniline dye is used. This adds to the color of the ferric salt so that ultimately a good black ink is produced. Some substance like phenol is added to prevent molding.

An inferior and cheaper ink can be made by dissolving aniline dyes in water. The most common one used is nigrosine. It is apt to settle out of solution when evaporation takes place and so be unpleasant. It does not produce so good a black nor so permanent a color. This ink washes out readily from any fabric.

Red writing inks are made by dissolving aniline dyes such as carmine and eosin in water. These two are removed by chlorine, but as a rule red ink is difficult to remove because one does not recognize the dye by simply looking at it. Occasionally cochineal, a dye made from



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Marble Quarry, Carrara, Italy

a tropical insect, is used as a basis for red inks. It is more easily removed than the preceding.

Copying inks like hectograph ink and typewriter ink are like the above-mentioned with the addition of sugar, gum, and deliquescent salts like magnesium or calcium chloride. They are usually very concentrated. Printers' ink and India ink are made from finely divided carbon as a basis. For printers' ink the carbon is mixed with varnish or linseed oil. This ink is of course not attacked by any chemical, though sometimes it may be partly removed by adding some solvent for the varnish or oil. India ink is of the same nature though not so concentrated as printers' ink.

Indelible inks may be either vegetable or mineral. Some plants like the cashew nut produce juices which mixed with lime water or alkali gives an indelible ink. Such inks may possibly be removed by chlorine. Silver nitrate makes the basis of the mineral inks. Silver nitrate and gum arabic and a small quantity of lampblack dissolved in ammonia, when treated with the heat of a flatiron, produce a fixed black due to the reduction of the silver nitrate within the fiber of the material.

BUILDING MATERIALS

Calcium Carbonate, $CaCO_3$.

Calcium carbonate or marble (limestone) is found in nature in large quantities. Often marble is pure white, but it may be colored pink, or gray, or green, due to presence of some other metallic compound. Limestone is a non-crystalline form and not so pure. Marble is found in large quantities in Italy, in Vermont, and in Nevada.

Because of its hardness, as well as its beautiful coloring, it finds a great use as a building material. From the chemical point of view it is useful as a source of carbon dioxide, the gas used for charging various mineral waters. When used for this purpose it is in the form of a fine dust, the waste building material.

Precipitated calcium carbonate, a very fine powdery form, is used as a basis for many tooth powders and pastes. As whiting it finds a wide use in cleaning metals of their tarnishes.

EXPERIMENT 52

Purpose.

To make lime from calcium carbonate.

Directions.

Put several pieces of marble into an ignition tube. Close the tube with a one-holed rubber stopper in which is a glass tube passing into limewater. Heat gradually until the full heat of Bunsen burner is used.

When the limewater has decidedly changed, disconnect. When cool add a little water to contents of tube and test with litmus paper.

Questions. What familiar substance is given off when marble is heated? The residue is lime, calcium oxide. How does its properties differ from marble?

Calcium oxide, CaO.

Calcium oxide or lime is made commercially by the same process as that studied in the last experiment, heating or calcining calcium carbonate. Lime formed in this manner is known as quicklime, and is a grayish white powdery substance. It combines readily with water. With the evolution of much heat, and if a moderate amount of water is added, the result is a fine powdery substance, known as slaked lime, or calcium hydroxide, $Ca(OH)_2$. Lime is very extensively used in the making of mortar, cements, plaster and concrete. Slaked lime is the foundation of whitewash, so extensively used as a paint. Slaked lime mixed with water is known as milk of lime if there is lime enough present to give the appearance of milk. In this form it is used to soften water, and in many commercial processes. If slaked lime is dissolved in water and filtered, the filtrate is the lime water so frequently used in babies' food to furnish bone-building material, and as a mouth wash to neutralize the acid action and so prevent decay.

Calcium sulphate, $CaSO_4$.

Calcium sulphate occurs in nature as the mineral gypsum $CaSO_42H_2O$. When this is heated until about seventy-five per cent of its water of crystallization is driven off the resulting substance is known as "plaster of Paris"; when water is again added the crystals take up the water, thus becoming harder and becoming "set." As "plaster of Paris," calcium sulphate is used for making casts and reproductions, interior decorative work, and as a cementing material for glass and metal ware. "Stucco" is made by mixing the plaster with a glue solution.

Cements.

Cements are double salts of aluminum and calcium silicates which combine chemically with water to form a very hard mass. Portland cement is made from clay, or aluminum silicate, and limestone heated together. Cements do not need carbon dioxide to make them harden as does mortar, nor does it need to be dried to harden like plaster of Paris. Therefore they may be used under water. Cements are fast taking the place of natural building materials because of their cheapness of production and their durability.

PAINTS, STAINS, VARNISHES, SHELLACS

Paints are intimate mixtures of pigments, or coloring substances, and oil, usually linseed oil. Linseed oil is oxidized by air, and so is hardened or dried. The pigment is of the most importance from the chemists' point of view. Lead compounds furnish the basis for many paints, lead carbonate and hydroxide giving "white lead," lead chromate, a yellow pigment, "chrome yellow," lead oxide giving "red lead." Iron compounds, especially the oxides, give ocher, Sienna, umber and rouge according to the different oxide or the amount of water it possesses. Copper acetate and carbonate are sometimes used for green pigments. Black paints have as their basis finely divided carbon. Barium, calcium, strontium and magnesium carbonates and sulphates are also used as white pigments. Zinc oxide is also coming rapidly into use as a white paint, and serves its purpose much better than the lead paints which interact with any sulphur compounds that may be in the air forming a black lead sulphide.

Zinc sulphide formed in the same way is white. Vermilion is a mercury sulphide.

Many of these pigments have been superseded or replaced by the colors made from coal tar. They are usually used as "Lakes," a lake being a compound of an inorganic substance with an organic coloring substance. They are classed as acid, basic or mordant pigments.

Varnishes are made by dissolving resins in turpentine, an oil from certain trees. Great care must be taken in its preparation in order that the varnish shall dry evenly and smoothly.

Shellac is a resinous substance exuded by certain insects as they feed upon certain trees. The color, as in the case of all resins, depends upon the plants from which it comes.

COMMON MEDICINAL SUBSTANCES

Magnesium sulphate, $MgSO_4$.

Magnesium sulphate or Epsom salts is made from the residue of the carnallite solution used in the making of potash compounds. It is used for medicinal purposes because of its action on the intestinal walls.

Aluminum sulphate, $Al_2(SO_4)_3$.

Aluminum sulphate may be made from (1) China clay, which is a silicate of aluminum, by treating with sulphuric acid; (2) by treating bauxite, an oxide of aluminum, with sulphuric acid; or (3) from the alumina (Al_2O_3) obtained in the cryolite soda process. The second method gives much the purest sulphate, and since the bauxite is more easily decomposed than the clay, the process is a much more usable one.

If a solution of aluminum sulphate is treated with a solution of sodium or potassium sulphate and the water allowed to evaporate large crystals of a double salt separate out. These crystals contain twenty-four molecules of water, and are known as alum. Other sulphates than sodium or potassium may be used, but in all cases the term "alum" is applied. Formerly alum was used extensively as a mordant in the dyeing business, as a sizing for paper, in tanning skins, and for purifying of water. Since the manufacture of aluminum sulphate results in so satisfactory a product free from iron the double salt has been largely replaced by it. Alum is sometimes used for its astringent action under the name of burnt alum. This is alum without its water of crystallization.

EXPERIMENT 53

Purpose.

To use aluminum sulphate, or alum, as a mordant.

Directions.

a. To a solution of aluminum sulphate, or alum, in a test-tube, add a little ammonium hydroxide. Add a little litmus to the tube, shake, and note appearance of the litmus. Allow it to settle.

b. Put some alizarine and cochineal into two evaporating dishes. Put some ammonium hydroxide and alum solution into two dishes.

c. Take strips of white cotton cloth which has no starch or dressing in it. Dip a piece into the alum, then into the ammonia water, then into the dye, alizarine. Repeat with cochineal. Boil for about three minutes. Note color changes.

d. Dip two pieces directly into the alizarine and cochineal without putting them into the alum and ammonia solutions. Boil or three minutes. Note color changes.

e. Remove the cloth from the dishes and wash each set carefully. Note color changes.

Questions. In *a* a "lake" is formed. What difference in color between the pieces of cloth dipped and not dipped? Which process gives the "fastest" color? (The alum and ammonia act as a mordant, making color set better.)

Mercurous chloride, HgCl.

Mercurous chloride or calomel is a white crystalline solid, a salt which is frequently used as a medicine. It is insoluble in water. It is prepared by combining mercury and chlorine.

Mercuric chloride, $HgCl_2$.

Mercuric chloride, which chemically may be considered a close relative to mercurous chloride, has properties quite different. It is an extremely poisonous substance. Because of this it finds an extensive use as a disinfectant; a solution of one part in a thousand parts of water being sufficient to cause death to living organisms. It is a white crystalline solid and unlike calomel soluble in water. It is prepared by the interaction of chlorine and mercury.

Iodine.

Iodine used to be obtained largely from seaweed, but now the more important method consists of getting iodine from the Chili saltpeter mother liquors.

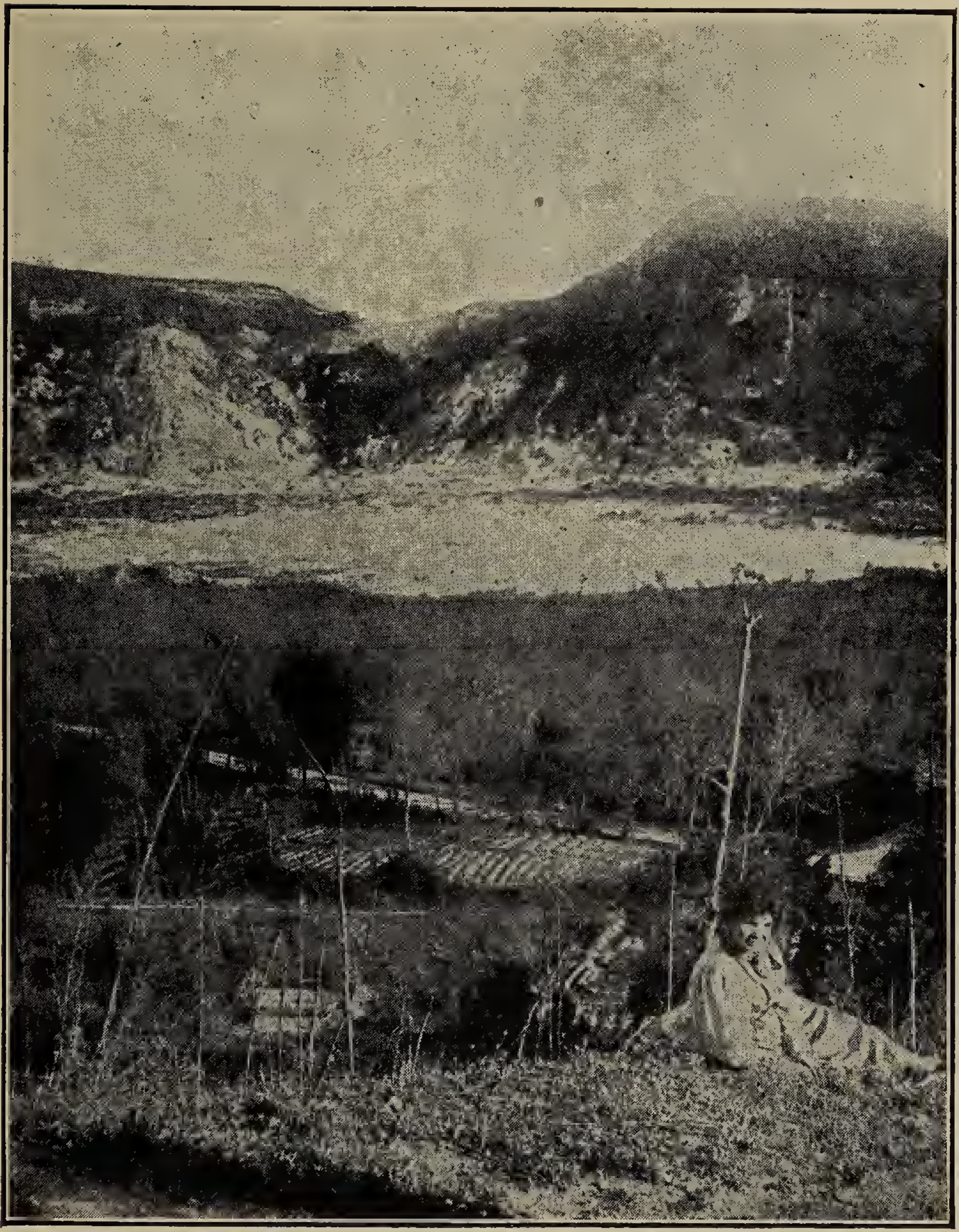
Seaweeds are burned leaving only the ashes. These ashes contain sodium and potassium iodides as well as other salts. These other salts are crystallized out by repeated evaporation until only the iodides are left. Manganese dioxide and sulphuric acid are then added to the iodides, which sets the iodine free. This is the method frequently used in the laboratory. Sometimes chlorine is passed into this mother liquor, which sets free the iodine.

The mother liquor from the manufacture of Chili saltpeter also contains iodine in the form of iodates. The iodine is precipitated out of this solution of iodates by shaking it with sulphur dioxide and water. The iodine can then be filtered, washed and dried, and purified by sublimation.

Sulphur.

Sulphur is found free in volcanic regions, especially in Sicily, California and Louisiana. In Louisiana wells are driven into the deposit, and hot water under pressure is forced down into the sulphur bed. The melted sulphur is forced to the surface through a pipe inside the well. The sulphur so obtained is quite pure. In Sicily piles of sulphur mixed with the earthy materials are heated, thereby freeing the sulphur from the rock material. Much sulphur is lost however by this method.

Crude sulphur is purified by heating and allowing



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White Sulphur Beds. Crater of Solfatara, Italy

the vapors to pass into a receiving chamber. If the temperature of this chamber is not above 110° C. the vapors condense at once to a fine powder and are sold under the name of flowers of sulphur. If the temperature rises above this temperature the vapors condense as a liquid which may be drawn off into molds. This is known as the roll brimstone. Sulphur thus prepared is crystalline, but under the right conditions sulphur may be amorphous. Sulphur like carbon is an allotropic element.

Sulphur is also found combined as sulphides and sulphates but is never prepared from its compounds. Sulphur found as iron sulphide or pyrites is the basis for the preparation of sulphuric acid, the most important sulphur compound. Sulphur as such is used in many fungicides and insecticides, such as Bordeaux mixture, and for the preparation of carbon bisulphide. The refined sulphur is used for making certain kinds of matches, for black gunpowder, and for vulcanizing rubber.

Boric acid, H_3BO_3.

Boric or boracic acid occurs free in volcanic regions but much is made from borax. The acid is a white crystalline solid somewhat soluble in cold water. It is used as a flux in enamels and glazes, and as a mild antiseptic, and as a preservative. Medically it is used as an eyewash.

COMMON SUBSTANCES

Ammonia, NH_3.

Ammonia is produced as a by-product during the destructive distillation of soft coal. The ammonia is set free and collected as ammoniacal liquor. This is sometimes distilled to collect the ammonia, and sometimes passed directly into sulphuric acid forming the sulphate. Ammonia is itself a gas, very soluble in water, and in solution is known as ammonia water or chemically ammonium hydroxide NH_4OH. Ammonia water has many uses in the house as a cleansing agent.

Ammonia is used extensively in the manufacture of artificial ice. This is because the gas is liquefied easily, and in the liquid condition evaporates easily, thus absorbing heat from anything that may be near.

Ammonia is the end product in the destructive distillation of animal matter, such as the waste material from slaughter houses. It is made in the animal body during the metabolic processes. The odor of ammonia is often noticeable in the refuse of stables and in the waste products of the human body. Because of this ammonia has sometimes been made by the destructive distillation of animal hair, hides, scrap leather and so on.

Since it is a base it combines with acids forming salts. The name "ammonium" is given to these salts because the ammonia NH_3 combines with part of the hydrogen of the water forming NH_4, ammonium. This has never been isolated but acts like a metal, giving its name to the salt. Ammonium salts find much use as fertilizers because of the large amounts of nitrogen present, especially ammonium nitrate.

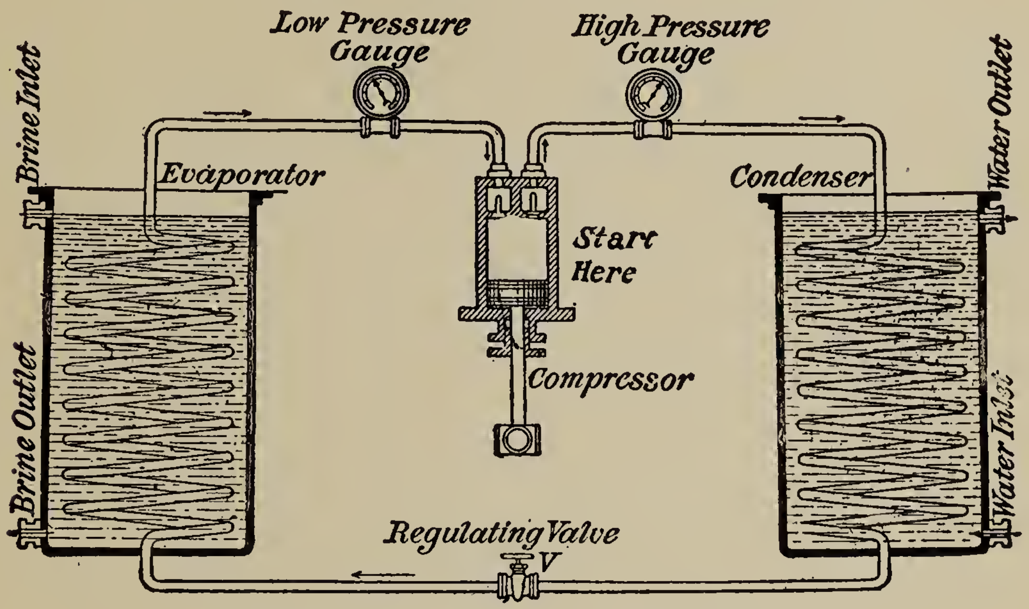

Diagram to Show the Working of an Artificial Ice Plant

The ammonia gas is in the compressor, is condensed to liquid in pipes leading to condenser, goes through *V* to the brine tank, where it turns back to gas as pressure is removed at *V*. During this last change any water surrounded by the brine is cooled.

EXPERIMENT 54

Purpose.

To make ammonia.

Directions.

a. Mix a little ammonium chloride and slaked lime in a mortar and grind very vigorously for a few seconds. Smell carefully of the mixture.

b. Using an ignition tube as generator set up apparatus to collect gas by upward dry displacement. Put into generator a mixture of ammonium chloride and lime and heat gently. Hold a moistened pink litmus paper at mouth of collecting bottle to show when it is full. Collect three bottles.

c. Observe physical properties of gas in one bottle.

d. Invert carefully a bottle of gas in a dish of water. Let it stand until decided action has occurred. Test water with dry litmus.

e. Test combustibility of gas in second bottle.

f. Hold the stopper of hydrochloric acid bottle at mouth of the third bottle.

Questions. Give physical and chemical properties of gas. Write reaction in symbols.

Sodium carbonate, Na_2CO_3.

Sodium carbonate or soda ash is one of the most important compounds of sodium. It is made by two different processes, each one producing about one half of the world's supply. The oldest method (1791) is known as the LeBlanc process. This involves the treating of sodium sulphate with carbon, giving sodium sulphide and carbon dioxide. The sulphide is treated with limestone, whereby the carbonate of soda is formed together with calcium sulphide. Hydrochloric acid and bleaching powder are produced as by-products of this process and for this reason the LeBlanc process is a profitable one to the manufacturer. The second and more recent (1863) process for making soda ash is known as the Solvay process. In this sodium chloride is treated with ammonia and carbon dioxide, which results in the formation of sodium acid carbonate. This is subjected to heat, which gives the normal sodium carbonate. This process gives no by-products, but is commercially successful because the ammonium chloride formed can be used as the source of ammonia for the further production of sodium carbonate. Soda ash as sold usually is a powdery solid with a slight red tinge and containing black specks. It is soluble in water giving alkaline reaction. It is used in making certain kinds of glass, and in many industrial processes for various reasons. If soda ash is dissolved in warm water, the sediment filtered off, and the liquor allowed to evaporate, large crystals of nearly pure soda or sal soda or washing soda separate out. These contain much water and are more bulky. This is not so soluble as soda ash, but is preferred by some manufacturers.

Sodium bicarbonate, $NaHCO_3$.

Sodium bicarbonate or baking soda is made during the Solvay process and is sold directly as such by the manufacturers. It may also be made by exposing the crystals of sal soda to an atmosphere of carbon dioxide. It is a white solid with a not unpleasant taste, soluble in water. It is used in cooking, as its name implies, and together with acid potassium tartrate makes what is known as baking powder. It is also used as a starting point in the manufacture of carbon dioxide for the making of soda water.

EXPERIMENT 55

Purpose.

To make baking powder, and test its action.

Directions.

a. Make separate solutions of sodium bicarbonate, $NaHCO_3$, and acid potassium tartrate or cream of tartar, $KHC_4H_4O_6$. Test both with litmus paper.

b. Make an intimate mixture of these two salts, using two and one-fourth times as much tartrate as bicarbonate. This should give practically no residue of either, according to molecular weight. Put in a test-tube. Add water, and arrange to let the gas formed pass into limewater.

c. After action has been completed, evaporate the contents of the tube to dryness.

Questions. With respect to the action on litmus, what kind of salts are there? What is the agent producing chemical action? What are the evidences of chemical action? In what does the value of baking powder lie? What is left as an end product? In the ordinary use of baking powder what becomes of this residue?

Write reaction in symbols.

Iron Sulphate, $FeSO_4$.

Iron sulphate, green vitriol or copperas is made as a by-product in many industries, such as in the distilling of pyrites for sulphur, and the weathering of copper sulphide ores. Dissolving scrap iron in the "sludge acid" of the petroleum refining process, and in the "pickling" liquors used for cleansing purposes in foundries, will give iron sulphate. Iron sulphate is formed from any of the above solutions by crystallization. The crystals are green in color, efflorescing quickly in air, and becoming covered with a brownish white powder. It is used as a mordant in dyeing, in the making of fuming sulphuric acid, for disinfecting purposes, for the making of inks, Prussian blue, and in various pigments.

Sodium tetraborate, $Na_2B_4O_7$.

Sodium tetraborate, commonly known as borax, is found native in Thibet, Ceylon and California. Borax mixed with sodium carbonate, sulphate and chloride is found in Nevada. In this country commercial borax is produced by dissolving the crude borax and recrystallizing. By crystallizing from a solution of 22° Bé. and allowing it to cool to 27° C., the common or prismatic form is produced. These are large monoclinic crystals which effloresce in air. They melt in their water of crystallization when heated, swelling into a spongy mass which finally becomes transparent and hard. By crystallizing from a solution which weighs 30° Bé. and cooled to 56° C., octohedral borax is formed. These crystals are permanent in dry air, but absorb moisture on exposure and pass into the prismatic kind. It fuses without swelling, and is therefore preferable for many purposes. Powdered borax may be a finely divided form of either variety.

Borax finds much use as a cleansing agent in the household, partly because of its alkaline quality and partly because a solution of it dissolves metallic oxides. It is used as a flux in welding processes, in enamels and glazes, in soaps, especially those used with hard water, as a meat

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Salt Fields, Le Croisic, France
Children collecting salt for the stacks.

preservative, as a mordant, and for medicinal purposes. Borax treated with either hydrochloric or sulphuric acid is decomposed into boric or boracic acid.

Sodium chloride, NaCl.

Sodium chloride or common salt is more frequently used than any other salt. It was the substance which gave the name " salt " to that large class of compounds studied in Chapter VII. It is obtained in large quantities from sea water, rock salt or from brines obtained from springs, lakes or wells. In our own country it is made from sea water on the coast of California. Sea water contains about two and a half per cent of salt. Large quantities of water from the Pacific Ocean are allowed to run into enclosures made by dams and the water evaporated by the heat of the sun. In Russia it is made by freezing the water and removing the ice, which leaves brine from which the salt is obtained by evaporation. The salt thus made is very coarse and usually absorbs water very easily, due to the presence of magnesium chloride. Such a method cannot be used on the Atlantic coast because the condition of the atmosphere prevents rapid evaporation.

Rock salt or halite is found in large deposits, and sometimes quite free from impurities, in England, Germany, Spain and Louisiana. It is mined in the United States in New York, Kansas, California, Utah and Louisiana.

Most of the salt produced in this country is made from natural brines found in New York, Michigan, Kansas and Ohio. Smaller quantities are produced in West Virginia, Utah, Texas and Pennsylvania. Sometimes, as in the case of the deposit near Syracuse, the brine is obtained by boring wells and raising the brine to the surface by pumps. The water is then evaporated either by the sun or by the use of fuel.

The properties of salt are familiar to everybody: its white color, agreeable taste, crystalline form, solubility in water. The use of salt with the daily food, and with ice as a freezing mixture, needs no comment here.

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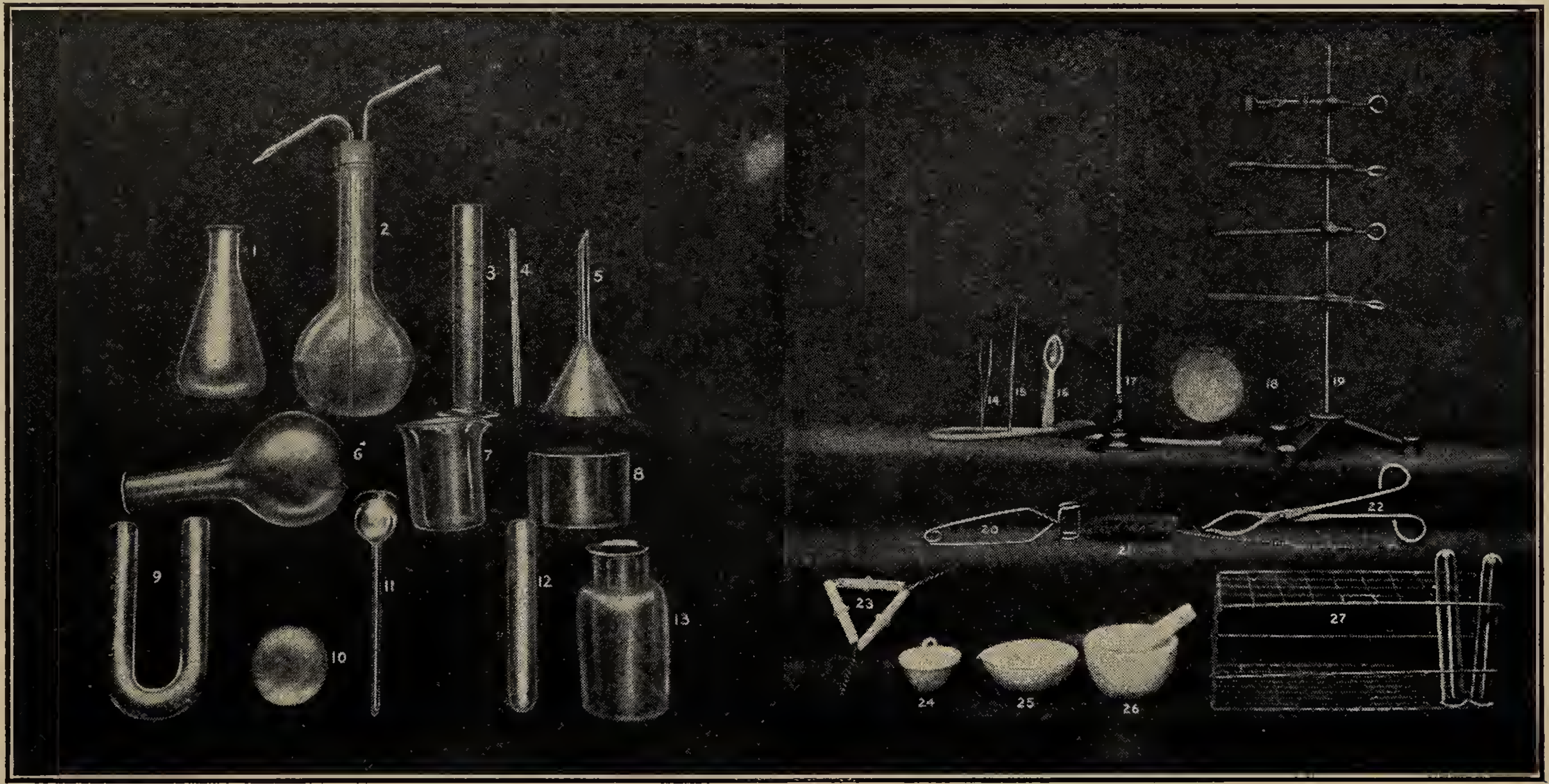

1. Erlenmeyer flask
2. Wash bottle
3. Graduate
4. Stirring rod
5. Funnel
6. Round-bottom flask
7. Beaker
8. Crystallizing dish
9. U-tube
10. Watch glass
11. Thistle tube
12. Test-tube
13. Bottle
14. Forceps
15. File
16. Horn spoon
17. Bunsen burner and tubing
18. Asbestos wire gauze
19. Ring stand, rings, and clamp
20. Test-tube holder
21. Test-tube brush
22. Crucible tongs
23. Clay triangle
24. Crucible and cover
25. Evaporating dish
26. Mortar and pestle
27. Test-tube rack

APPENDIX A

DIRECTIONS FOR LABORATORY WORK

On Performing Experiments. — The complete directions, including Questions at the end, should be read first, to gain a general idea of the experiment as a whole. The apparatus and materials should then be assembled and arranged if necessary, and the number and purpose of the experiment, the drawing, apparatus, materials and data headings written in the notebook. The experiment should then be performed, notes of happenings and of weights jotted down on scrap paper as they occur, and the description, calculations and conclusion written during or after the experiment. The writing should always be completed immediately after the work is done.

On Describing Experiments. — The best results will be obtained if a uniform plan for recording experiments be adopted by all members of a class. All such plans should contain certain essential features, although they may vary in details of arrangement. These features are as follows:

1. The number and purpose of the experiment.

2. A drawing of apparatus, if this would be helpful in describing the work or in recalling the experiment to memory later in the course. These drawings should be simple outlines of what would appear were a cross section made of the apparatus.[1]

3. A list of apparatus and materials used, unless the experiment is so simple that this is unnecessary.

4. Data in tabulated form where required.

5. A description, in exact, concise language, of what is done and what occurs. This may include descriptions of substances, answers to questions, and other points

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[1] See the drawings that accompany many of the experiments. The use of a stencil, such as that sold by L. E. Knott, Cambridge, Mass., will save much time and insure accuracy.

suggested in the directions. It should constantly be borne in mind while writing this that the drawing and descriptions are the only source of information for the work after it is completed, and they should both be made with the view to recalling the experiment accurately and vividly to mind later.

6. A conclusion made of answers to the Questions at the end of the experiment. These will sometimes summarize the information gained, and sometimes emphasize the importance of one special feature of the experiment which it is desirable to dwell on particularly.

On Describing Substances. — Substances are described by enumerating their qualities or properties. These include the physical properties of weight, color, odor, luster, hardness, toughness or elasticity. Such a description of sulphur, for instance, would state that it is a light yellow substance of medium weight, with no odor except when burning (sulphur dioxide), slight luster, and that it crumbles when broken or cut. Accurate descriptions are a very necessary feature of every record of chemical experiments.

On Using the Bunsen Burner. — The burner permits two kinds of flame, — the yellow or illuminating flame and the colorless or heating flame. When the holes at the base of the tube are open, air is admitted with the gas, combustion is more rapid and the flame more concentrated and hotter. The hottest part of the flame is just above the tip of the inner cone. The inner cone itself is cool and no burning occurs there. By varying the amount of opening at the base a flame of varying intensity may be obtained.

On Heating Apparatus. — Thin glass ware may be heated very hot, provided the beaker or flask be protected from the flame by a wire gauze, provided it be perfectly dry outside, and provided no direct heat be applied above the contents. In heating a substance in a test-tube, keep the tube moving slightly in the flame to heat all sides evenly, and do not heat beyond the substance in it. In evaporating to dryness use porcelain dishes and be careful not to heat too long or too hot. Porcelain crucibles will stand the hottest flame directly applied.

On Weighing. — The object to be weighed is placed on the left hand pan. If it is a chemical reagent, first place a piece of paper on the pan and balance this with another piece or with sand on the right hand pan. Never handle chemicals with the fingers; use a spoon, spatula or small piece of paper. Set weights carefully on the right hand pan, being careful not to jar the balance if it is a delicate one. Always handle small aluminum weights with forceps, as moisture from the hand changes their weight. Always record the weight at the balance before the separate weights are returned to the box. Write the name of the article or substance and its weight together on a piece of paper. This is later transferred to the notebook. Leave the balance at rest, and in a clean condition, and return all weights to their places.

APPENDIX B

METRIC SYSTEM

Length.

10 millimeters (mm.) = 1 centimeter (cm.).
10 cm. = 1 decimeter (dm.).
10 dm. = 1 meter (m.).

The unit of length is the meter, 39.37 in., or a little more than a yard. The common unit in laboratory work is the centimeter, about $\frac{3}{8}$ of an inch.

Surface.

100 sq. mm. = 1 sq. cm.
100 sq. cm. = 1 sq. dm.
100 sq. dm. = 1 sq. m.

Volume.

1000 cu. mm. = 1 cu. cm. (c. c.).
1000 c. c. = 1 cu. dm. or 1 liter (l.).

The common units in laboratory work are the cubic centimeter and the liter. A test-tube holds between 20 and 30 c. c. ordinarily. A liter is a little more than a liquid quart.

10 milligrams (mg.) = 1 centigram (cg.).
10 cg. = 1 decigram (dg.).
10 dg. = 1 gram (gm.).

The gram is the common unit, about 28 grams making an ounce avoirdupois.

One cubic centimeter of water at 4° C. weighs one gram.

THERMOMETERS

The thermometer usually used in the laboratory is the Centigrade thermometer. The freezing point of water is marked 0°, and the boiling point of water is 100°.

The thermometer used in the home is the Fahrenheit. The freezing point is marked 32°, and the boiling point 212°.

Since there are 100° difference between boiling and freezing points on the Centigrade and 180° on the Fahrenheit, a Centigrade degree is larger than the Fahrenheit, and 0° on the Fahrenheit is 32° below the freezing point.

It is sometimes convenient to change from a reading on one thermometer to a reading on the other. To change from C.° to F.° take 9/5 of the number of C.° and add 32. To make the opposite change subtract 32 from the number of F°. and take 5/9 of the result.

$$\text{F.}^\circ = 9/5\ \text{C.} + 32$$
$$\text{C.}^\circ = 5/9\ (\text{C.} - 32)$$

DISPLACEMENT SERIES OF METALS

Potassium
Sodium
Calcium
Magnesium
Aluminum
Manganese
Zinc
Chromium
Iron
Nickel
Tin
Lead
Hydrogen
Copper
Bismuth
Antimony
Mercury
Silver
Platinum
Gold

APPENDIX C

TABLE OF FOOD VALUES

In the following table two sets of values will often be found, "edible portion" and "as purchased." From the numbers under "edible portion" may be estimated the value of food as eaten; from those under "as purchased" some comparison may be made between cost and value received.

To use the table intelligently some general considerations should be held in mind. The large percent of water in all foods, even in those called dried, will at once strike the attention. It is evident that too concentrated a diet is undesirable. It should not be forgotten that the calorific value is not the sole test of a food's quality, as the fiber, the acids and salts of vegetables and fruits are as important as are energy-producing properties. At the same time, certain expensive foods of low calorific value, such as quail, oysters, terrapin, asparagus, celery and lettuce, might well be replaced by other fresh and less expensive kinds where economy must be practiced.

Among foods of animal origin, notice particularly the low percent of carbohydrates and the high protein and fat values. Where the percent of fat is high, notice the increase in the calorific values, as in the dairy products, such as butter and cheese.

Among vegetable foods notice, especially in flours, the high value for carbohydrates. Interesting comparisons may be made between individual foods. The difference between sago and tapioca, for instance, in the matter of protein will be surprising to most people. Lima and string beans, too, present an interesting comparison. Nuts are seen to be a concentrated food of high calorific value, some being rich in fat, like the walnut, some in carbohydrates, like the chestnut.

THE CHEMICAL COMPOSITION OF AMERICAN FOOD MATERIALS[1]

ANIMAL FOOD

FOOD MATERIALS	Refuse	Water	Protein by difference	Fat	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.[3]
Beef, cooked							
Cut not given, boiled, as purchased	...	38.1	26.1	34.9	...	0.9	2805
Roast, as purchased, average[2]	...	48.2	22.3	28.6	...	1.3	1620
Pressed, as purchased	...	44.1	26.7	27.7	...	1.5	1610
Round steak, fat removed, as purchased, average	...	63.0	27.5	7.7	...	1.8	840
Sirloin steak, baked, as purchased	...	63.7	24.7	10.2	...	1.4	875
Loin steak, tenderloin, broiled, edible portion, average	...	54.8	23.6	20.4	...	1.2	1300
Soup stock	...	89.1	5.8	1.5	...	3.6	170
Sweetbreads, uncooked, as purchased	...	70.9	15.4	12.1	...	1.6	825
Suet, as purchased, average	...	13.7	4.2	81.8	...	0.3	3540
Beef, canned							
Corned beef, average	...	51.8	25.5	18.7	...	4.0	1280
Dried beef, as purchased	...	44.8	38.6	5.4	...	11.2	960
Tongue, whole, as purchased, average	...	51.3	21.5	23.2	...	4.0	1340
Tripe, as purchased, average	...	74.6	16.4	8.5	...	0.5	670
Veal kidneys, as purchased, average	...	75.8	16.5	6.4	...	1.3	585
Veal liver, as purchased, average	...	73.0	20.4	5.3	...	1.3	575

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[1] Compiled from Bulletin No. 28 (Revised) U. S. Dept. of Agriculture. The Chemical Composition of American Food Materials. W. O. Atwater and A. P. Bryant, Supt. of Documents Gov. Printing Office. 10c.

[2] Average means the average of several analyses of the same kind of food, except in case of cookies, cakes, etc., where the average for different varieties is given.

[3] For a valuable table of 100 Caloric portions in common measures, such as cups, slices, etc., see M. S. Rose, "Feeding the Family," Appendix, Tables II and III.

ANIMAL FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein by difference	Fat	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.
Lamb, cooked							
Chops, broiled, average, edible portion	...	47.6	21.2	29.9	...	1.3	1665
Leg, roast	...	67.1	19.4	12.7	...	0.8	900
Mutton, cooked							
Leg, roast, edible portion, average	...	50.9	25.3	22.6	...	1.2	1420
Pork							
Ham, fresh, average							
edible portion	...	50.1	15.6	33.4	...	0.9	1700
as purchased	10.3	45.1	14.1	29.7	...	0.8	1520
Ham, smoked, average							
edible portion	...	39.8	16.7	38.8	...	4.7	1945
as purchased	12.2	35.8	14.6	33.2	...	4.2	1670
Shoulder, smoked, average							
edible portion	...	37.6	15.3	41.0	...	6.1	2020
as purchased	18.9	30.7	12.4	33.0	...	5.0	1625
Pigs' feet, pickled, average							
edible portion	...	68.2	16.1	14.8	...	0.9	930
as purchased	35.5	44.6	10.0	9.3	...	0.6	585
Salt pork, lean ends, average							
edible portion	...	19.9	7.3	67.1	...	5.7	2985
as purchased	11.2	17.6	6.5	59.6	...	5.1	2655
Bacon, smoked, average							
edible portion	...	20.2	9.9	64.8	...	5.1	2930
as purchased	8.7	18.4	9.0	59.4	...	4.5	2685
Ribs, cooked, as purchased	...	33.6	26.6	37.6	...	2.2	2050
Steak, cooked, as purchased	...	33.2	19.9	45.4	...	1.5	2285
Ham, deviled, as purchased, average	...	44.1	18.5	34.1	...	3.3	1790
Sausage							
Arles							
edible portion	...	17.2	24.9	50.6	...	7.3	2635
as purchased	5.2	16.3	23.6	48.0	...	6.9	2495
Bologna, average							
edible portion	...	60.0	18.4	17.6	0.3	3.7	1095
as purchased	3.3	55.2	18.0	19.7	...	3.8	1170
Frankfort, as purchased, average	...	57.2	19.7	18.6	1.1	3.4	1170

ANIMAL FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein by difference	Ash	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Poultry and Game							
Capon, cooked							
edible portion	...	59.9	27.3	11.5	...	1.3	985
as purchased	10.4	52.6	24.5	10.3	...	1.2	885
Chicken fricasseed, edible portion	...	67.5	17.6	11.5	2.4	1.0	855
Turkey, roast, edible portion	...	52.0	28.4	18.4	...	1.2	1295
Turkey, roast, light and dark meat and stuffing, edible portion	...	65.0	17.1	10.8	5.5	1.6	870
Chicken, sandwich, canned, as purchased	...	46.9	20.5	30.0	...	2.6	1655
Turkey, sandwich, canned, as purchased	...	47.4	20.7	29.2	...	2.7	1615
Quail, canned, as purchased	...	66.9	21.8	8.0	1.7	1.6	775
Fish							
Blue fish, cooked, edible portion	...	68.2	26.1	4.5	...	1.2	670
Spanish mackerel, broiled							
edible portion	...	68.9	23.2	6.5	...	1.4	715
as purchased	7.9	63.5	21.4	5.9	...	1.3	655
Cod, salt, average							
edible portion	..	53.5	21.5	0.3	...	24.7	410
as purchased	24.9	40.2	16.0	0.4	...	18.5	315
Cod, salt, " boneless "							
edible portion, average	...	55.0	25.7	0.3	...	19.0	490
as purchased	1.6	54.8	28.6	0.3	...	14.7	545
Herring, smoked							
edible portion	...	34.6	36.4	15.8	...	13.2	1355
as purchased	44.4	19.2	20.2	8.8	...	7.4	750
Mackerel, salt, entrails removed							
edible portion	...	42.2	22.0	22.6	...	13.2	1345
as purchased	22.9	32.5	17.0	17.4	...	10.2	1035
Salmon, canned, average							
edible portion	...	63.5	21.8	12.1	...	2.6	915
as purchased	14.2	56.8	19.5	7.5	...	2.0	680
Sardines, canned							
edible portion, average	...	52.3	22.4	19.7	...	5.6	1260
as purchased	5.0	53.6	24.0	12.1	...	5.3	950

ANIMAL FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein by difference	Ash	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.
Sturgeon, caviare, pressed Russian, as purchased	...	38.1	30.0	19.7	7.6	4.6	1530
Trout, brook, canned							
edible portion	..	68.4	22.8	6.1	...	3.7	670
as purchased	3.5	66.1	20.9	5.9	...	3.6	650
Tunny, canned, as purchased	...	72.7	21.5	4.1	...	1.7	575
Frogs' legs, average							
edible portion	...	83.7	15.1	0.2	...	1.0	295
as purchased	32.0	56.9	10.3	0.1	...	0.7	200
Shell Fish, etc., fresh							
Clams, round, removed from shell, as purchased	...	80.8	10.6	1.1	5.2	2.3	340
Crabs, hardshell, whole							
edible portion	...	77.1	16.6	2.0	1.2	3.1	415
as purchased	52.4	36.7	7.9	0.9	0.6	1.5	195
Lobsters, whole, average							
edible portion	...	79.2	16.4	1.8	0.4	2.2	390
as purchased	61.7	30.7	5.9	0.7	0.2	0.8	140
Oysters, solids, average, as purchased	...	88.3	6.0	1.3	3.3	1.1	230
Scallops, average, as purchased	...	80.3	14.8	0.1	3.4	1.4	345
Terrapin							
edible portion	...	74.5	21.0	3.5	...	1.0	545
as purchased	75.4	18.3	5.2	0.9	...	0.2	135
Eggs							
Hens', uncooked							
edible portion, average	...	73.7	14.8	10.5	...	1.0	720
as purchased	11.2	65.5	13.1	9.3	...	0.9	635
Hens', boiled							
edible portion, average	...	73.2	14.0	12.0	...	0.8	765
as purchased	11.2	65.0	12.4	10.7	...	0.7	680
Hens', boiled, whites, average	...	86.2	13.0	0.2	...	0.6	250
Hens', boiled, yolks, average	...	49.5	16.1	33.3	...	1.1	1705
Dairy Products, etc.							
Butter, as purchased	...	11.0	1.0	85.0	...	3.0	3605
Buttermilk, as purchased	...	91.0	3.0	0.5	4.8	0.7	165

ANIMAL FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein by difference	Fat	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Cheese							
American, pale, as purchased	...	31.6	28.8	35.9	0.3	3.4	2055
Cottage, as purchased, average	...	72.0	20.9	1.0	4.3	1.8	510
Limburger, as purchased	...	42.1	23.0	29.4	0.4	5.1	1675
Neufchâtel, as purchased, average	...	50.0	18.7	27.4	1.5	2.4	1530
Pineapple, as purchased, average	...	23.0	29.9	38.9	2.6	5.6	2245
Roquefort, as purchased	...	39.3	22.6	29.5	1.8	6.8	1700
Swiss, as purchased, average	...	31.4	27.6	34.9	1.3	4.8	2010
Koumiss, as purchased, average	...	89.3	2.8	2.1	5.4	0.4	240
Milk,							
condensed, sweetened, average, as purchased	...	26.9	8.8	8.3	5.4	11.9	1520
unsweetened, "evaporated cream," as purchased, average	...	68.2	9.6	9.3	11.2	1.7	780
whole, as purchased	...	87.0	3.3	4.0	5.0	0.7	325
Miscellaneous							
Gelatin, as purchased, average	...	13.6	84.2	0.1	...	2.1	1705
Lard, refined, as purchased	...	...	...	100.0	...	...	4220
Cottolene, as purchased	...	...	...	100.0	...	...	4220
Oleomargarine, as purchased	...	9.5	1.2	83.0	...	6.3	3525
Beef juice, as purchased	...	93.0	4.9	0.6	...	1.5	115

VEGETABLE FOOD

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.
Flour, Meals, etc.							
Barley, pearled, average	...	11.5	8.5	1.1	77.8	1.1	1650
Buckwheat flour, average	...	13.6	6.4	1.2	77.9	0.9	1620
Corn flour, average	...	12.6	7.1	1.3	78.4	0.6	1645
Corn meal, granular, average	...	12.5	9.2	1.9	75.4	1.0	1655
Popcorn, average	...	4.3	10.7	5.0	78.7	1.3	1875
Hominy, average	...	11.8	8.3	0.6	79.0	0.3	1650
cooked	...	79.3	2.2	0.2	17.8	0.5	380
Oatmeal, average	...	7.3	16.1	7.2	67.5	1.9	1860
boiled	...	84.5	2.8	0.5	11.5	0.7	285
Rolled oats, average	...	7.7	16.7	7.3	66.2	2.1	1850
Rice, average	...	12.3	8.0	0.3	79.0	0.4	1630
boiled, average	...	72.5	2.8	0.1	24.4	0.2	510
Rye flour, average	...	12.9	6.8	0.9	78.7	0.7	1630
Wheat flour, California fine, average	...	13.8	7.9	1.4	76.4	0.5	1625
entire wheat, average	...	11.4	13.8	1.9	71.9	1.0	1675
patent roller process, family and straight grade, average	...	12.8	10.8	1.1	74.8	0.5	1640
high grade, average	...	12.4	11.2	1.0	74.9	0.5	1645
low grade, average	...	12.0	14.0	1.9	71.2	0.9	1665
Wheat preparations							
Cracked and crushed wheat, average	...	10.1	11.1	1.7	75.5	1.6	1685
Farina, average	...	10.9	11.0	1.4	76.3	0.4	1685
Flaked wheat, average	...	8.7	13.4	1.4	74.3	2.2	1690
Shredded wheat, average	...	8.1	10.5	1.4	77.9	2.1	1700
Macaroni, average	...	10.3	13.4	0.9	74.1	1.3	1665
cooked	...	78.4	3.0	1.5	15.8	1.3	415
Spaghetti, average	...	10.6	12.1	0.4	76.3	0.6	1660
Bread, Crackers, Pastry, etc.							
Bread, brown, as purchased, average	...	43.6	5.4	1.8	47.1	2.1	1050
Johnnycake, as purchased, average	...	38.9	7.9	4.7	46.3	2.2	1205

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Bread, wheat							
Graham, as purchased, average	...	35.7	8.9	1.8	52.1	1.5	1210
Biscuit, home-made, as purchased, average	...	32.9	8.7	2.6	55.3	0.5	13 0
Rolls, as purchased, average	...	29.2	8.9	4.1	56.7	1.1	1395
White bread, as purchased, average	...	35.3	9.2	1.3	53.1	1.1	1215
Whole wheat bread, as purchased, average	...	38.4	9.7	0.9	49.7	1.3	1140
Zwieback, as purchased, average	...	5.8	9.8	9.9	73.5	1.0	1970
Crackers, as purchased, average	...	6.8	10.7	8.8	71.9	1.8	1905
Cake, as purchased, except fruit, average	...	19.9	6.3	9.0	63.3	1.5	1675
Cookies, as purchased, average	...	8.1	7.0	9.7	73.7	1.5	1910
Wafers, as purchased, average	...	6.6	7.6	11.6	72.9	1.3	1985
Doughnuts, as purchased, average	...	18.3	6.7	21.0	53.1	0.9	2000
Pie							
apple, as purchased, average	...	42.5	3.1	9.8	42.8	1.8	1270
custard, as purchased	...	62.4	4.2	6.3	26.1	1.0	830
lemon, as purchased	...	47.4	3.6	10.1	37.4	1.5	1190
mince, as purchased, average	...	41.3	5.8	12.3	38.1	2.5	1335
squash, as purchased	...	64.2	4.4	8.4	21.7	1.3	840
Pudding							
rice custard, as purchased	...	59.4	4.0	4.6	31.4	0.6	825
tapioca, as purchased, average	...	64.5	3.3	3.2	28.2	0.8	720
Sugars, Starches, etc.							
Candy, as purchased	...	...	...	...	96.0	...	1785
Marshmallows, 27% starch and flour	...	5.6	...	...	57.4	1.1	....
Caramels, 32.2% starch and flour	...	3.3	...	...	52.7	1.4	....
Chocolate creams, 15.4% starch and flour	...	3.8	...	...	72.1	0.5	....
Honey, as purchased, average	...	18.2	0.4	...	81.2	0.2	1520

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.
Molasses, cane, as purchased, average	...	25.1	2.4	...	69.3	3.2	1920
Cornstarch, as purchased	...	...	...	...	90.0	...	1675
Sago, as purchased	...	12.2	9.0	0.4	78.1	0.3	1635
Tapioca, as purchased, average	...	11.4	0.4	0.1	88.0	0.1	1650
Sugar, brown, as purchased	...	...	...	...	95.0	...	1765
granulated, as purchased	...	...	...	...	100.0	...	1860
maple, as purchased, average	...	...	...	...	82.8	...	1540
powdered, as purchased	...	...	...	...	100.0	...	1860
Syrup, maple, as purchased, average	...	...	...	...	71.4	...	1330
Vegetables							
Artichokes, as purchased, average	...	79.5	2.6	0.2	16.7	1.0	365
Asparagus, as purchased							
fresh, average	...	94.0	1.8	0.2	3.3	0.7	105
cooked	...	91.6	2.1	3.3	2.2	0.8	220
Beans,							
dried, as purchased, average	...	12.6	22.5	1.8	59.6	3.5	1605
Lima, dried, as purchased, average	...	10.4	18.1	1.5	65.9	4.1	1625
Lima, fresh,							
edible portions	...	68.5	7.1	0.7	22.0	1.7	570
as purchased	55.0	30.8	3.2	0.3	9.9	0.8	255
string, cooked, edible portion	...	95.3	0.8	1.1	1.9	0.9	95
string, fresh							
edible portion, average	...	89.2	2.3	0.3	7.4	0.8	195
as purchased	7.0	83.0	2.1	0.3	6.9	0.7	180
Beets, fresh,							
cooked, edible portion	...	88.6	2.3	0.1	7.4	1.6	185
edible portion, average	...	87.5	1.6	0.1	9.7	1.1	215
as purchased	20.0	70.0	1.3	0.1	7.7	0.9	170
Cabbage,							
edible portion, average	...	91.5	1.6	0.3	5.6	1.0	145
as purchased	15.0	77.7	1.4	0.2	4.8	0.9	125
Carrots, fresh							
edible portion, average	...	88.2	1.1	0.4	9.3	1.0	210
as purchased	20.0	70.6	0.9	0.2	7.4	0.9	160

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	Cals.
Cauliflower, as purchased, average	...	92.3	1.8	0.5	4.7	0.7	140
Celery							
edible portion, average	...	94.5	1.1	0.1	3.3	1.0	85
as purchased	20.0	75.6	0.9	0.1	2.6	0.8	70
Corn, green							
edible portion, average	...	75.4	3.1	1.1	19.7	0.7	470
as purchased	61.0	29.4	1.2	0.4	7.7	0.3	180
Cucumbers,							
edible portion, average	...	95.4	0.8	0.2	3.1	0.5	80
as purchased	15.0	81.1	0.7	0.2	2.6	0.4	70
Eggplant, edible portion	...	92.9	1.2	0.3	5.1	0.5	130
Greens,							
beet, cooked, as purchased	...	89.5	2.2	3.4	3.2	1.7	245
dandelion, as purchased	...	81.4	2.4	1.0	10.6	4.6	285
Kohl-rabi, edible portion, average	...	91.1	2.0	0.1	5.5	1.3	145
Leeks,							
edible portion	...	91.8	1.2	0.5	5.8	0.7	150
as purchased	15.0	78.0	1.0	0.4	5.0	0.6	130
Lentils, dried, as purchased, average	...	8.4	25.7	1.0	59.2	5.7	1620
Lettuce,							
edible portion, average	...	94.7	1.2	0.3	2.9	0.9	90
as purchased	15.0	80.5	1.0	0.2	2.5	0.8	75
Mushrooms, as purchased, average	...	88.1	3.5	0.4	6.8	1.2	210
Onions, fresh,							
edible portion, average	...	87.6	1.6	0.3	9.9	0.6	225
as purchased	10.0	78.9	1.4	0.3	8.9	0.5	205
cooked, prepared, as purchased	...	91.2	1.2	1.8	4.9	0.9	190
green,							
edible portion average,	...	87.1	1.0	0.1	11.2	0.6	230
as purchased	51.0	42.6	0.5	0.1	5.5	0.3	115
Parsnips,							
edible portion, average	...	83.0	1.6	0.5	13.5	1.4	300
as purchased	20.0	66.4	1.3	0.4	10.8	1.1	240
Peas,							
dried, as purchased, average	...	9.5	24.6	1.0	62.0	2.9	1655

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
Peas, green,	%	%	%	%	%	%	Cals.
edible portion, average	...	74.6	7.0	0.5	16.9	1.0	465
as purchased	45.0	40.8	3.6	0.2	9.8	0.6	255
green, cooked, as purchased	...	73.8	6.7	3.4	14.6	1.5	540
Potatoes, raw or fresh							
edible portion, average	...	78.3	2.2	0.1	18.4	1.0	385
as purchased	20.0	62.6	1.8	0.1	14.7	0.8	310
cooked, average							
boiled, as purchased	...	75.5	2.5	0.1	20.9	1.0	440
chips, as purchased	...	2.2	6.8	39.8	46.7	4.5	2675
mashed and creamed, as purchased	...	75.1	2.6	3.0	17.8	1.5	505
sweet, raw or fresh							
edible portion, average	...	69.0	1.8	0.7	27.4	1.1	570
as purchased	20.0	55.2	1.4	0.6	21.9	0.9	460
sweet, cooked and prepared, as purchased	...	51.9	3.0	2.1	42.1	0.9	925
Pumpkins,							
edible portion, average	...	93.1	1.0	0.1	5.2	0.6	120
as purchased	50.0	46.5	0.5	0.1	2.6	0.3	60
Radishes,							
edible portion, average	...	91.8	1.3	0.1	5.8	1.0	135
as purchased	30.0	64.3	0.9	0.1	4.0	0.7	95
Rhubarb,							
edible portion, average	...	94.4	0.6	0.7	3.6	0.7	105
as purchased	40.0	56.6	0.4	0.4	2.2	0.4	65
Spinach,							
fresh, as purchased, average	...	92.3	2.1	0.3	3.2	2.1	110
cooked, as purchased	...	89.8	2.1	4.1	2.6	1.4	260
Squash,							
edible portion, average	...	88.3	1.4	0.5	9.0	0.8	215
as purchased	50.0	44.2	0.7	0.2	4.5	0.4	105
Tomatoes,							
fresh, as purchased, average	...	94.3	0.9	0.4	3.9	0.5	105
dried, as purchased	...	7.3	12.9	8.1	62.3	9.4	1740
Turnips							
edible portion, average	...	89.6	1.3	0.2	8.1	0.8	185
as purchased	30.0	62.7	0.9	0.1	5.7	0.6	125
Pickles, Condiments, etc.							
Catsup, tomato, as purchased, average	...	82.8	1.5	0.2	12.3	3.2	265

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Horseradish, as purchased, average	...	86.1	1.4	0.2	10.5	1.5	230
Olives, green,							
edible portion	...	58.0	1.1	27.6	11.6	1.7	1400
as purchased	27.0	42.3	0.8	20.2	8.5	1.2	1025
Olives, ripe,							
edible portion	...	64.7	1.7	25.9	4.3	3.4	1205
as purchased	19.0	52.4	1.4	21.0	3.5	2.7	975
Pickles, mixed, as purchased	...	93.8	1.1	0.4	4.0	0.7	110
Fruits, Berries, etc., fresh							
Apples							
edible portion, average	...	84.6	0.4	0.5	14.2	0.3	290
as purchased	25.0	63.3	0.3	0.3	10.8	0.3	220
Bananas, yellow							
edible portion, average	...	75.3	1.3	0.6	22.0	0.8	460
as purchased	35.0	48.9	0.8	0.4	14.3	0.6	300
Blackberries, as purchased, average	...	86.3	1.3	1.0	10.9	0.5	270
Cherries,							
edible portion, average	...	80.9	1.0	0.8	16.7	0.6	365
as purchased	5.0	76.8	0.9	0.8	15.9	0.6	345
Cranberries, as purchased, average	...	88.9	0.4	0.6	9.9	0.2	215
Currants, as purchased	...	85.0	1.5	...	12.8	0.7	265
Figs, fresh, as purchased, average	...	79.1	1.5	...	18.8	0.6	380
Grapes,							
edible portions, average	...	77.4	1.3	1.6	19.2	0.5	450
as purchased	25.0	58.0	1.0	1.2	14.4	0.4	335
Huckleberries, edible portion	...	81.9	0.6	0.6	16.6	0.3	345
Lemons							
edible portion, average	...	89.3	1.0	0.7	8.5	0.5	205
as purchased	30.0	62.5	0.7	0.5	5.9	0.4	145
Muskmelons							
edible portion	...	89.5	0.6	...	9.3	0.6	185
as purchased	50.0	44.8	0.3	...	4.6	0.3	90
Nectarines							
edible portion	...	82.9	0.6	...	15.9	0.6	305
as purchased	6.6	77.4	0.6	...	14.8	0.6	285

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	*%*	*%*	*%*	*%*	*%*	*Cals.*
Oranges							
edible portion, average	...	86.9	0.8	0.2	11.6	0.5	240
as purchased	27.0	63.4	0.6	0.1	8.5	0.4	170
Peaches,							
edible portion, average	...	89.4	0.7	0.1	9.4	0.4	190
as purchased	18.0	73.3	0.5	0.1	7.7	0.3	155
Pears,							
edible portion, average	...	84.4	0.6	0.5	14.1	0.4	295
as purchased	10.0	76.0	0.5	0.4	12.7	0.4	260
Pineapple, edible portion	...	89.3	0.4	0.3	9.7	0.3	200
Plums,							
edible portion, average	...	78.4	1.0	...	20.1	0.5	395
as purchased	5.0	74.5	0.9	...	19.1	0.5	370
Pomegranates, edible portion, average	...	76.8	1.5	1.6	19.5	0.6	460
Prunes,							
edible portion, average	...	79.6	0.9	...	18.9	0.6	370
as purchased	5.8	75.6	0.7	...	17.4	0.5	335
Raspberries,							
red, as purchased	...	85.8	1.0	...	12.6	0.6	255
black, edible portion, average	...	84.1	1.7	1.0	12.6	0.6	310
Strawberries,							
edible portion, average	...	90.4	1.0	0.6	7.4	0.6	180
as purchased	5.0	85.9	0.9	0.6	7.0	0.6	175
Watermelons,							
edible portion, average	...	92.4	0.4	0.2	6.7	0.3	140
as purchased	59.4	37.5	0.2	0.1	2.7	0.1	60
Fruits, etc., dried							
Apples, as purchased, average	...	28.1	1.6	2.2	66.1	2.0	1350
Apricots, as purchased, average	...	29.4	4.7	1.0	62.5	2.4	1290
Citron, as purchased, average	...	19.0	0.5	1.5	78.1	0.9	1525
Currants, as purchased, average	...	17.2	2.4	1.7	74.2	4.5	1495
Dates,							
edible portion, average	...	15.4	2.1	2.8	78.4	1.3	1615
as purchased	10.0	13.8	1.9	2.5	70.6	1.2	1450
Figs, as purchased, average	...	18.8	4.3	0.3	74.2	2.4	1475
Pears, as purchased	...	16.5	2.8	5.4	72.9	2.4	1635

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Prunes,							
edible portion, average	...	22.3	2.1	...	73.3	2.3	1400
as purchased	15.0	19.0	1.8	...	62.2	2.0	1190
Raisins,							
edible portion, average	...	14.6	2.6	3.3	76.1	3.4	1605
as purchased	10.0	13.1	2.3	3.0	68.5	3.1	1445
Raspberries, as purchased	...	8.1	7.3	1.8	80.2	2.6	1705
Fruits, etc., canned, and Jellies, Preserves, etc.							
Apple sauce, as purchased	...	61.1	0.2	0.8	37.2	0.7	730
Figs, stewed, as purchased	...	56.5	1.2	0.3	40.9	1.1	785
Grape butter, as purchased	...	36.7	1.2	0.1	58.5	3.5	1115
Marmalade (orange peel)							
as purchased	...	14.5	0.6	0.1	84.5	0.3	1585
Pineapples, as purchased	...	61.8	0.4	0.7	36.4	0.7	715
Prune sauce, as purchased	...	76.6	0.5	0.1	22.3	0.5	460
Nuts							
Almonds							
edible portion, average	...	4.8	21.0	54.9	17.3	2.0	3030
as purchased	45.0	2.7	11.5	30.2	9.5	1.1	1660
Butternuts,							
edible portion	...	4.4	27.9	61.2	3.5	2.9	3165
as purchased	86.4	0.6	3.8	8.3	0.5	0.4	430
Chestnuts, fresh							
edible portion, average	...	45.0	6.2	5.4	42.1	1.3	1125
as purchased	16.0	37.8	5.2	4.5	35.4	1.1	945
Chestnuts, dried,							
edible portion, average	...	5.9	10.7	7.0	74.2	2.2	1875
as purchased	24.0	4.5	8.1	5.3	56.4	1.7	1425
Cocoanuts,							
edible portion	...	14.1	5.7	50.6	27.9	1.7	2760
as purchased	48.8	7.2	2.9	25.9	14.3	0.9	1413
Cocoanut without milk,							
as purchased	37.3	8.9	3.6	31.7	17.5	1.0	1730
Cocoanut, prepared, as purchased, average	...	3.5	6.3	57.4	31.5	1.3	3125
Filberts,							
edible portion	...	3.7	15.6	65.3	13.0	2.4	3290
as purchased	52.1	1.8	7.5	31.3	6.2	1.1	1575

VEGETABLE FOOD — *continued*

FOOD MATERIALS	Refuse	Water	Protein	Fat	Total carbohydrates (incl. fiber)	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Hickory nuts,							
edible portion	...	3.7	15.4	67.4	11.4	2.1	3345
as purchased	62.2	1.4	5.8	25.5	4.3	0.8	1265
Peanuts							
edible portion, average	...	9.2	25.8	38.6	24.4	2.0	2560
as purchased	24.5	6.9	19.5	29.1	18.5	1.5	1935
Peanut butter, as purchased	...	2.1	29.3	46.5	17.1	5.0	2825
Pecans, polished,							
edible portion	...	3.0	11.0	71.2	13.3	1.5	3455
as purchased	53.2	1.4	5.2	33.3	6.2	0.7	1620
Pecans, unpolished,							
edible portion	...	2.7	9.6	70.5	15.3	1.9	3435
as purchased	46.3	1.5	5.1	37.9	8.2	1.0	1846
Walnuts, California,							
edible portion	...	2.5	18.4	64.4	13.0	1.7	3300
as purchased	73.1	0.7	4.9	17.3	3.5	0.5	885
Miscellaneous							
Chocolate, as purchased, average	...	5.9	12.9	48.7	30.3	2.2	2860
Cocoa, as purchased, average	...	4.6	21.6	28.9	37.7	7.2	2320
Cereal coffee infusion (1 part boiled in 20 parts water)	...	98.2	0.2	...	1.4	0.2	30

ANIMAL AND VEGETABLE FOODS

FOOD MATERIAL	Refuse	Water	Protein N. × 6.25	Fat	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Soups, Home-made							
Bean soup, as purchased	...	84.3	3.2	1.4	9.4	1.7	295
Chicken soup, as purchased	...	84.3	10.5	0.8	2.4	2.0	275
Clam chowder, as purchased, average	...	88.7	1.8	0.8	6.7	2.0	195

ANIMAL AND VEGETABLE FOODS — *continued*

FOOD MATERIAL	Refuse	Water	Protein N. × 6.25	Fat	Total carbohydrates	Ash	Fuel value per pound
	%	%	%	%	%	%	*Cals.*
Soups, Canned							
Asparagus, cream of, as purchased	...	87.4	2.5	3.2	5.5	1.4	285
Bouillon, as purchased, average	...	96.6	2.2	0.1	0.2	0.9	50
Celery, cream of, as purchased	...	88.6	2.1	2.8	5.0	1.5	250
Chicken gumbo, as purchased, average	...	89.2	3.8	0.9	4.7	1.4	195
Chicken soup, as purchased, average	...	93.8	3.6	0.1	1.5	1.0	100
Consommé, as purchased	...	96.0	2.5	...	0.4	1.1	55
Corn, cream of, as purchased	...	86.8	2.5	1.9	7.8	1.0	270
Mock turtle, as purchased, average	...	89.8	5.2	0.9	2.8	1.3	185
Mulligatawny, as purchased, average	...	89.3	3.7	0.1	5.7	1.2	180
Oxtail,							
edible portion, average	...	88.8	4.0	1.3	4.3	1.6	210
as purchased	1.8	87.8	3.8	0.5	4.2	1.9	170
Pea soup, as purchased, average	...	86.9	3.6	0.7	7.6	1.2	235
Pea, cream of green, as purchased	...	87.7	2.6	2.7	5.7	1.3	270
Tomato soup, as purchased, average	...	90.0	1.8	1.1	5.6	1.5	185
Turtle, green, as purchased	...	86.6	6.1	1.9	3.9	1.5	265
Vegetable, as purchased	...	95.7	2.9	...	0.5	0.9	65
Miscellaneous							
Hash, as purchased	...	80.3	6.0	1.9	9.4	2.4	365
"Infants and Invalids' Foods" as purchased, average	...	6.0	12.7	3.3	76.2	1.8	1795
Mincemeat, commercial as purchased, average	...	27.7	6.7	1.4	60.2	4.0	1305
homemade as purchased, average	...	54.4	4.8	6.7	32.1	2.0	970
Salad, ham, as purchased	...	69.4	15.4	7.6	5.6	2.0	710
Sandwich,							
egg, as purchased	...	41.4	9.6	12.7	34.5	1.8	1355
chicken, as purchased	...	48.5	12.3	5.4	32.1	1.7	1055

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Name	Symbol	Atomic Weight O = 16	Valence	Name	Symbol	Atomic Weight	ence
Aluminum	Al.	27.1	III	Mc			VI
Antimony	Sb.	120.2	III, V	Ne			
Argon	A.	39.88		Ne			
Arsenic	As.	74.96	III, V	*Nic*			II, III
Barium	Ba.	137.37	II	Nit	Nt.	222.4	
Bismuth	Bi.	208.	III, V	*Nitrogen*	N.	14.01	III, V
Boron	B.	11.	III	Osmium	Os.	190.9	III
Bromine	Br.	79.92	I	*Oxygen*	O.	16.	II
Cadmium	Cd.	112.4	II	Palladium	Pd.	106.7	I, II, IV
Caesium	Cs.	132.81	I	*Phosphorus*	P.	31.04	III, V
Calcium	Ca.	40.07	II	*Platinum*	Pt.	195.2	IV
Carbon	C.	12.	IV	*Potassium*	K.	39.1	I
Cerium	Ce.	140.25		Praseodymium	Pr.	140.6	
Chlorine	Cl.	35.46	I	*Radium*	Ra.	226.4	II
Chromium	Cr.	52.	II, III, VI	Rhodium	Rh.	102.9	III
Cobalt	Co.	58.97	II, III	Rubidium	Rb.	85.45	I
Columbium	Cb.	93.5		Ruthenium	Ru.	101.7	III
Copper	Cu.	63.57	I, II	Samarium	Sa.	150.4	
Dysprosium	Dy.	162.5		Scandium	Sc.	44.1	III
Erbium	Er.	167.7		Selenium	Se.	79.2	IV
Europium	Eu.	152.		*Silicon*	Si.	28.3	IV
Fluorine	F.	19.	I	*Silver*	Ag.	107.88	I
Gadolinium	Gd.	157.3		*Sodium*	Na.	23.	I
Gallium	Ga.	69.9	III	*Strontium*	Sr.	87.63	II
Germanium	Ge.	72.5		*Sulphur*	S.	32.07	II, IV, VI
Glucinum	Gl.	9.1	II	Tantalum	Ta.	181.5	V
Gold	Au.	197.2	I, III	Tellurium	Te.	127.5	II
Helium	He.	3.99		Terbium	Tb.	159.2	
Holmium	Ho.	163.5		Thallium	Tl.	204.	I, III
Hydrogen	H.	1.008	I	Thorium	Th.	232.4	
Indium	In.	114.8	III	Thulium	Tm.	168.5	
Iodine	I.	126.92	I	*Tin*	Sn.	119.	II, IV
Iridium	Ir.	193.1	III, IV	Titanium	Ti.	48.1	IV
Iron	Fe.	55.84	II, III	*Tungsten*	W.	184.	IV
Krypton	Kr.	82.92		Uranium	U.	238.5	I, VI, VIII
Lanthanum	La.	139.	III	*Vanadium*	V.	51.	V
Lead	Pb.	207.1	II, IV	Xenon	Xe.	130.2	
Lithium	Li.	6.94	I	Ytterbium	Yb.	172.	III
Lutecium	Lu.	174.		Yttrium	Yt.	89.	III
Magnesium	Mg.	24.32	I	*Zinc*	Zn.	65.37	II
Manganese	Mn.	54.93	II, IV	Zirconium	Zr.	90.6	
Mercury	Hg.	200.6	I, II				

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COMMON RADICALS

Ammonium	NH_4	I	Nitrite	NO_2	I
Carbonate	CO_3	II	Phosphate	PO_4	III
Hydroxyl	OH	I	Sulphate	SO_4	II
Nitrate	NO_3	I	Sulphite	SO_3	II